



International Correspondence Schools
Scranton, Pa.

Inorganic Chemistry

Laws of Matter and Energy Atomic and Molecular Weights
Solutions Ionization Physical Processes

PREPARED ESPECIALLY FOR HOME STUDY

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INORGANIC CHEMISTRY

(PART 2)

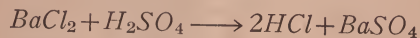
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Edition 1

CHEMICAL THEORY AND LAWS

LAWS OF MATTER AND ENERGY

1. **Law of Conservation of Mass.**—An analytical study of chemical changes shows that, if the substances involved in the chemical change are weighed before and after the change, there is no change or alteration in the quantity of matter. Consider, for example, the chemical change that takes place when barium chloride is acted upon by sulfuric acid:



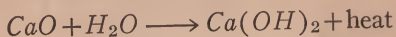
$$208.274 + 98.076 \longrightarrow 72.93 + 233.42$$

It has been determined by experiment that 208.274 parts by weight of BaCl_2 and 98.076 parts by weight of H_2SO_4 produce 72.93 parts by weight of 2HCl and 233.42 parts by weight of BaSO_4 ; that is, 306.35 parts by weight of the substances used form 306.35 parts by weight of the substances produced. In other words, the sum of the weights of substances used is equal to the sum of the weights of the products formed.

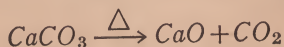
The fact that the weight of the products formed is equal to the weight of the substances used, is true of all chemical changes and leads to an important chemical law, the law of conservation of mass: *The mass of a system is not affected by any chemical change within the system.* This law may be stated differently, as follows: Whenever a change takes place in the composition of substances, the weight of matter after the change is the same as the weight of the matter involved before the change.

2. Conservation of Energy.—Closely allied to the law regarding the conservation of mass is the law relating to the conservation of energy. Like matter, energy can be neither created nor destroyed, but unlike matter, energy neither has weight nor occupies space. Energy may be defined as the capacity to do work. Heat, light, and electricity are different forms of energy, all of which have capacities for doing work; each of these forms of energy can bring about physical and chemical changes.

In all chemical changes, energy is either liberated or absorbed. For example, if calcium oxide, commonly known as quicklime, is mixed with water, a considerable quantity of energy in the form of heat is liberated and slaked lime is produced:



Similarly, energy in the form of heat must be applied to change calcium carbonate into calcium oxide and carbon dioxide:

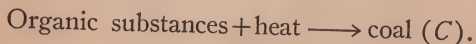


The various forms of energy can be converted into one another. For example, the heat under a boiler in an electric plant is used to generate steam, which in turn causes an engine to produce motion. The motion produced is used to run a dynamo, which produces electricity; if the current of electricity is passed through an incandescent lamp, light is produced and heat is liberated. It is seen that the form of energy, heat, changes to the other forms, such as motion, electricity, and light. If it were possible to produce these transformations in an apparatus from which no form of energy could escape, the quantity of energy at the beginning of the transformation would be found to be equal to the quantity of energy at the finish. Since neither a creation nor a destruction of energy has ever been observed, the following may be stated, as a law: *Energy cannot be created nor destroyed.*

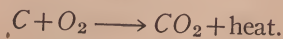
3. Chemical Energy.—Chemical energy is that form of energy which is liberated or absorbed when a chemical change takes place. Like all other forms of energy, chemical energy has the capacity to do work.

When calcium carbonate, $CaCO_3$, is decomposed into calcium oxide, CaO , and carbon dioxide, CO_2 , heat is absorbed; whereas on the other hand, when calcium oxide CaO , reacts with water to form slaked lime, $Ca(OH)_2$, heat is liberated. The heat that was liberated by the reaction between calcium oxide and water was due to the heat that was absorbed by the calcium oxide during its formation.

The heat energy liberated by the burning of coal may be accounted for in a similar manner. During the formation of coal, heat is absorbed:



Coal burns, or combines with oxygen, to form carbon dioxide and liberates heat:



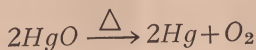
From the foregoing statement, it is apparent that chemical changes are always accompanied by some kind of energy transformation. Chemical energy may be changed to some other form of energy, and other forms of energy may be converted into chemical energy.

It should be noted here that neither the liberation nor the absorption of heat is an exclusive mark of chemical change. All physical changes are likewise accompanied by the liberation or absorption of energy. For example, when water evaporates it absorbs heat, and when liquids solidify they liberate heat.

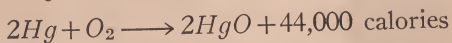
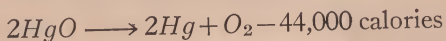
THERMOCHEMISTRY

4. General Remarks.—When substances react to form new compounds, heat is either liberated or absorbed. The branch of chemistry which deals with energy changes manifested by heat phenomena is called thermochemistry. This absorption or liberation of heat may not be apparent, nevertheless it may be measured by the use of a suitable apparatus, called a calorimeter (Greek, heat measurer). The unit of heat is the calorie, which may be defined as the quantity of heat required to raise the temperature of 1 gram of water 1 degree on the centigrade scale.

In the decomposition of mercuric oxide, HgO , for example, heat is absorbed in the actual disruptive process:



It has been calculated that, when 2 gram-molecules (that is, $2 \times$ molecular weight of HgO), or 433.22 grams, of mercuric oxide is decomposed, producing 2 gram-atoms of mercury (401.22 grams) and 1 gram-molecule of oxygen gas (32 grams), 44,000 calories of heat is absorbed. It also has been calculated that, when these same quantities of the separate elements combine to form mercuric oxide, the same quantity of heat is liberated. The above facts are expressed thus:

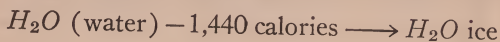
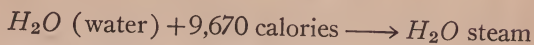


The foregoing reactions show that the system $[2Hg + O_2 \longrightarrow 2HgO]$ contains an excess of energy over the system $[2HgO \longrightarrow 2Hg + O_2]$ amounting to 44,000 calories. It must be clearly understood that nothing is known concerning the total energy in either of these systems. Though the amount of heat absorbed or liberated by a system can be measured, the total amount of heat in a system can not be estimated. The amount of heat energy which is absorbed or liberated in a system by chemical change is called free, or chemical, energy, while the non-transferable energy is the bound, or latent, energy. The chemical and latent energy compose the total energy of a system.

5. Thermochemical Equations.—The equations which represent a chemical change in matter accompanied by heat absorption or liberation are known as *thermochemical* equations. These are based on the principle of the conservation of energy in the same manner as the chemical or mass equation is based on the conservation of matter.

Before illustrating by examples the range of thermochemistry, it is necessary to make known the fact that thermal equations are not necessarily thermochemical equations. They may be physical, rather than chemical, changes accompanied by the absorption or liberation of heat.

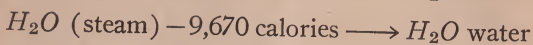
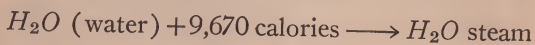
The physical transformation between water, ice, and steam will illustrate the foregoing statement:



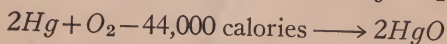
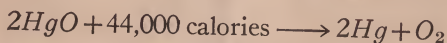
Thus, when water weighing 18.016 grams is converted into the same weight of steam, 9,670 calories of heat is absorbed, whereas, when this quantity of water is converted into the same weight of ice, 1,440 calories of heat is liberated.

Observation of the thermal equations mentioned shows that, in the case of a reversible change, whether it is physical or chemical, if the liberation of a certain quantity of heat accompanies the change in one direction, the absorption of the same quantity of heat must accompany the change when it is reversed.

For example, in the case of the physical change of water into steam,



The same is true of a reversible chemical change:



When 18.016 grams of water is converted into steam, 9,670 calories of heat is applied, and to condense the steam the exact quantity of heat, 9,670 calories, must be removed. The same is true of the chemical change of HgO ; 44,000 calories of heat is required to decompose 2 molecules of HgO , and the same quantity of heat is liberated when the change is reversed.

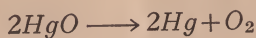
6. Endothermic and Exothermic Reactions.—The reactions in which heat is liberated are known as exothermic, while the reactions which absorb heat are known as endothermic. These terms are of great value when chemical changes involving a great quantity of heat energy are considered. The example in Art. 4 shows that, since the decomposition of HgO requires or absorbs 44,000 calories, it must be an endothermic reaction; whereas, when the elements mercury and oxygen combine to

form the mercuric oxide (HgO), 44,000 calories is liberated, and an exothermic reaction takes place.

The two reactions, one the formation of the oxide and the other the dissociation of the oxide, are equal and opposite in regard to energy and mass, since they are based on the principles of the conservation of energy and the conservation of mass. In the light of the foregoing information it is obvious that the heat of formation of mercuric oxide is +44,000 calories, that is, 44,000 calories of heat is liberated; whereas, the heat of decomposition is -44,000 calories, that is, 44,000 calories of heat is absorbed. The heat of formation may be defined as the amount of heat evolved or absorbed when 1 gram-molecular weight of a compound is formed. When heat is absorbed, the value found is negative.

LAWS OF CHEMICAL ACTION

7. Law of Definite Proportions.—In the decomposition of mercuric oxide (HgO), it is found that, for every 200.61 parts of mercury liberated, 16 parts of oxygen by weight are set free. By using the numbers which represent the atomic weights, a condensed statement can be formed to represent the same proportion by weight:



$$433.22 \text{ g.} \longrightarrow 401.22 \text{ g.} + 32 \text{ g.}$$

From experimental observation we know that mercury and oxygen combine to form mercuric oxide, and the proportion by weight in the combination is the same. For example, 401.22 grams of mercury will unite with 32 grams of oxygen to form 433.22 grams of mercuric oxide. Every sample of mercuric oxide, regardless of its method of formation, contains the same proportion by weight of the elements mercury and oxygen.

In like manner the compound magnesium oxide, MgO , which results from a union of magnesium, Mg , and oxygen, O , contains a definite proportion by weight of these two elements. The atomic weight of magnesium is 24.32 and that of oxygen is 16. The molecular weight of magnesium oxide is equal to the sum of the atomic weights of magnesium and oxygen, or $24.32 + 16 = 40.32$. In the decomposition of magnesium oxide,

24.32 parts by weight of magnesium are liberated and 16 parts by weight of oxygen are set free. Experimental investigations have shown that, regardless of the method of formation, magnesium oxide always contains these relative weights in definite proportion; that there is never any ratio between magnesium and oxygen other than 24.32:16.

The foregoing observations are found to apply to all chemical compounds, and are summed up in a general way in the law of definite proportions: *A chemical compound always contains the same constituents in the same proportion by weight.*

8. Law of Multiple Proportions.—The existence of two chlorides of tin offers a noteworthy fact, which appears more striking when it is discovered that in one compound the proportion of tin to chlorine is exactly twice as much as the proportion in the other compound. Tin and chlorine form two compounds: stannous chloride, SnCl_2 , and stannic chloride, SnCl_4 . By referring to a table of atomic weights, one is able to calculate the parts by weight of chlorine that unite with a fixed weight of tin. For instance, the atomic weight of tin, Sn , is 118.7 and that of chlorine, Cl , 35.457. There are $2 \times 35.457 = 70.914$ parts by weight of chlorine to every 118.7 parts by weight of tin in SnCl_2 ; likewise, there are $4 \times 35.457 = 141.828$ parts by weight of chlorine to every 118.7 parts by weight of tin in SnCl_4 . The parts by weight of chlorine in stannous chloride and in stannic chloride, combined with 118.7 parts by weight of tin, bear a ratio to each other of 70.914:141.828. When the terms of this ratio are reduced to their lowest values by dividing by 70.914, the ratio becomes 1:2. This ratio is a simple one because it is expressed in whole numbers.

There are also two oxides of mercury: mercurous oxide, Hg_2O , and mercuric oxide, HgO . The proportion of mercury to oxygen in mercurous oxide is 401.22:16; while in mercuric oxide the proportion of mercury to oxygen is 200.61:16. Therefore, the proportions by which mercury will unite with oxygen bear a ratio to each other of 401.22:200.61. When the terms of this ratio are reduced to simple whole numbers, the ratio becomes 2:1. No other proportions have been found to exist.

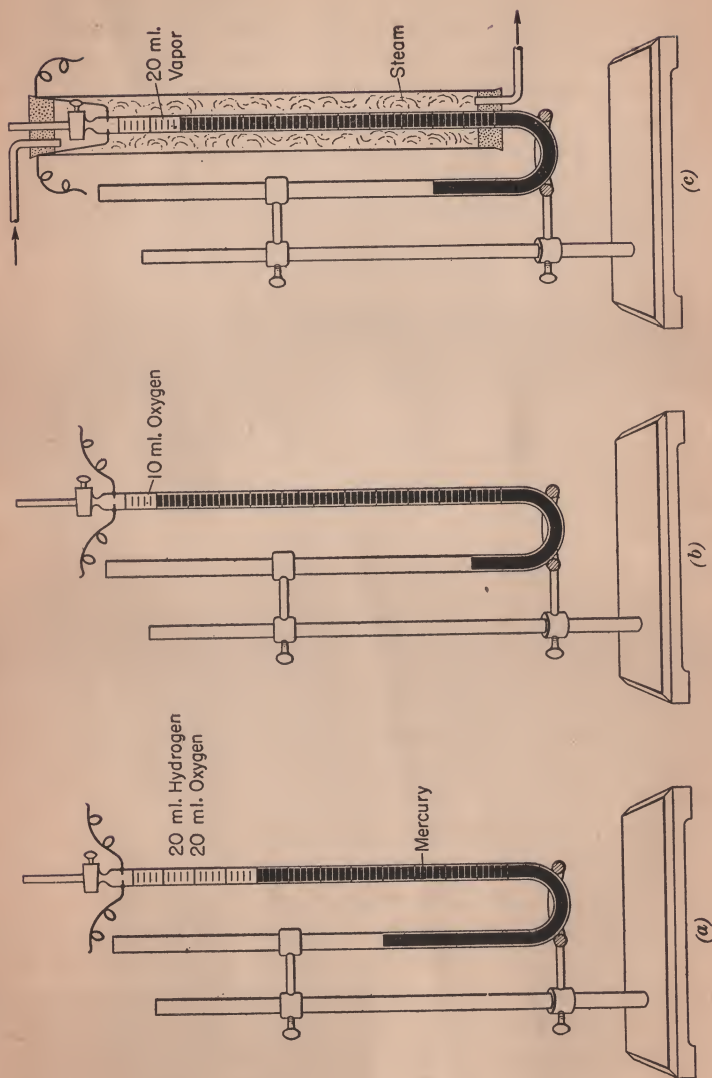


FIG. 1

The foregoing illustrations are examples of a discovery made by Dalton (1804) which he postulated as the law of multiple proportions: *When two elements combine in more proportions than one, forming two or more compounds, the weights of the element that combines with a fixed weight of the other bear a simple ratio to each other, expressed in small integral numbers.*

9. Law of Combining Weights.—The law of definite proportions creates an interest in the ratio of the weights of elements that combine to form a chemical compound. These combining weights are easily calculated by the analysis of pure compounds.

Oxygen, since it combines directly with most elements, is used as a standard. It is possible, by actual experiment, to determine what weight of each element will combine with a fixed weight of oxygen. The weight of the element that combines with a fixed weight of oxygen is known as the combining weight or equivalent weight of the element.

It has been noted by actual experiment that 8 parts of oxygen combine with 1.008 parts of hydrogen to form water. Further experiments with which 8 parts of oxygen combine with other substances show the following results:

- 8 parts of oxygen combine with 3 parts of carbon
- 8 parts of oxygen combine with 20.04 parts of calcium
- 8 parts of oxygen combine with 35.457 parts of chlorine
- 8 parts of oxygen combine with 39.096 parts of potassium
- 8 parts of oxygen combine with 12.16 parts of magnesium
- 8 parts of oxygen combine with 22.997 parts of sodium

Since 22.997 parts of sodium and 35.457 parts of chlorine each combine with 8 parts of oxygen, you would expect, since things equal to the same thing are equal to each other, 22.997 parts of sodium to combine with 35.457 parts of chlorine. Experiment has proved this to be true. In general, the combining capacity that a substance has with oxygen is a measure of its combining capacity when it combines with other substances to form compounds.

The combining weight of oxygen is given a value of 8 because 8 parts of oxygen combine with 1.008 parts of hydrogen, the

element with the lowest combining weight. The foregoing information is summarized in the law of combining or equivalent weights, which may be stated as follows: *The combining or equivalent weight of an element is the number of grams of that element which will combine with or displace 8 grams of oxygen or 1.008 grams of hydrogen.*

ATOMIC AND MOLECULAR WEIGHTS

10. Combining Ratio of Hydrogen and Oxygen.—The volume relationship of hydrogen and oxygen may be accurately determined by synthesizing water from the two elements, hydrogen and oxygen. By synthesis is meant the formation of a chemical compound from more elementary substances. The experiment of combining the two gases is carried on in a *eudiometer*, Fig. 1. A eudiometer is a graduated glass tube with two platinum wires fused through the glass at the top of the graduated glass tube in such a manner that a spark gap of 2 to 3 mm. exists between the ends of the platinum wires. To perform the experiment, a volume of exactly 20 ml. of oxygen and a like volume of hydrogen are placed in the graduated tube shown in (a), Fig. 1. Mercury is used to confine the two gases. When an electric spark is forced across the gap between the platinum wires by an induction coil, the two gases combine to form a small volume of water. Immediately the mercury rushes upward to fill the space formerly occupied by the gases. The upward travel stops at the 10 ml. mark as shown in (b), Fig. 1. Unless an excess of one of the gases is used, the mercury will strike the top of the tube with sufficient force to break it. In the example cited, the residual gas in (b) is found, upon examination, to be oxygen. The volume of oxygen measures exactly 10 ml. From the details of the experiment it is quite evident that the 20 ml. of hydrogen have combined with 10 ml. of oxygen. Therefore, the combining ratio of hydrogen and oxygen by volume is 2 to 1.

11. A modification of the foregoing experiment may be made by admitting exactly 20 ml. of hydrogen and 10 ml. of oxygen into the tube (c), Fig. 1. The tube containing the two

gases is then completely surrounded by an outer tube through which steam issues. The steam in this outer tube prevents the condensation of the steam produced in the eudiometer. When the spark is passed across the gap, all the hydrogen and the oxygen combine and the mercury rises to a point which indicates the presence of 20 ml. of some gas or vapor. Investigation shows the vapor to be pure steam. This experiment shows that 2 volumes of hydrogen combine with 1 volume of oxygen to form 2 volumes of water in the form of steam.

12. Gay-Lussac's Law of Combining Volumes.—The French chemist Gay-Lussac was so impressed with the fact that hydrogen and oxygen combine to form water in the simple and exact ratio of 2 to 1, that he studied the chemical actions between a number of gases. In every case he found that the volumes of the gases used and produced in any chemical change can always be represented by the ratio of small whole numbers.

The following are the results of some of the experiments that he conducted:

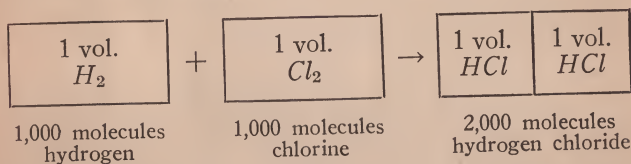
3 vol. hydrogen + 1 vol. nitrogen \longrightarrow 2 vol. ammonia
2 vol. carbon monoxide + 1 vol. oxygen \longrightarrow 2 vol. carbon dioxide
1 vol. hydrogen + 1 vol. chlorine \longrightarrow 2 vol. hydrogen chloride

Gay-Lussac summed up the results of his experiments in the form of a law, which may be stated as follows: *The volumes of the gases used and produced in any chemical change can always be represented by the ratio of small whole numbers.*

13. Avogadro's Hypothesis.—In 1811 Avogadro, an Italian physicist, proposed his famous hypothesis which accounts for the uniformity of results obtained by Gay-Lussac in his experiments. It is stated as follows: *Equal volumes of all gases contain the same number of molecules when subjected to the same conditions of temperature and pressure.* Stating the hypothesis differently it may be said that, under similar conditions of temperature and pressure, all gases have equal molecular concentrations; that is, a liter of oxygen, hydrogen, nitrogen, or any other gas, always contains the same number of molecules under similar conditions of temperature and pressure. On this assump-

tion Avogadro was able to explain satisfactorily the behavior of gases.

It has already been stated that 1 volume of hydrogen combines with 1 volume of chlorine to form 2 volumes of hydrogen chloride, HCl . If it is assumed that each volume of gas involved contains 1,000 molecules, then, according to Avogadro's hypothesis, the following relation can be expressed:



It can be readily seen from the foregoing expression that each molecule of hydrogen chloride contains 1 atom of hydrogen and 1 atom of chlorine. Therefore, there must be 2,000 hydrogen atoms and 2,000 chlorine atoms in 2,000 molecules of hydrogen chloride. Since these 2,000 atoms of hydrogen and 2,000 atoms of chlorine came from their respective 1,000 molecules of hydrogen and 1,000 molecules of chlorine, the hydrogen and the chlorine molecule must each contain 2 atoms.

Oxygen, hydrogen, and other elements that are gaseous at ordinary temperatures, usually have 2 atoms in a molecule. However, experiments show that this is not true of all elements. Phosphorus has 4 atoms to the molecule, while the number of atoms of sulfur vary from 2 to 8 per molecule, according to the temperature. The molecules of metals and of inert gases consist of single atoms.

14. Atomicity of Molecules.—The number of atoms in a molecule of an element is the atomicity of the element. According to whether the molecule contains 1, 2, 3, 4, or 6 atoms, it is respectively known as monatomic, diatomic, triatomic, tetraatomic, or hexatomic. The atomicities of many elements have been determined experimentally, whereas, in a few cases the atomicity is uncertain.

Table I shows the probable atomicity of many elements. The atomicities of some of the elements are known to be variable

TABLE I
NAMES, ATOMICITIES, AND MOLECULAR FORMULAS OF
CHEMICAL ELEMENTS

Name and Atomicity of Element		Molecular Formula	Name and Atomicity of Element		Molecular Formula
Monatomic	Carbon	<i>C</i>	Diatomic	Iodine	<i>I₂</i>
	Lithium	<i>Li</i>		Bromine	<i>Br₂</i>
	Potassium	<i>K</i>		Chlorine	<i>Cl₂</i>
	Sodium	<i>Na</i>		Fluorine	<i>F₂</i>
	Barium	<i>Ba</i>		Oxygen	<i>O₂</i>
	Strontium	<i>Sr</i>		Hydrogen	<i>H₂</i>
	Calcium	<i>Ca</i>		Nitrogen	<i>N₂</i>
	Magnesium	<i>Mg</i>		Bismuth	<i>Bi₂</i>
	Mercury	<i>Hg</i>		Antimony	<i>Sb₂</i>
	Zinc	<i>Zn</i>		Sulfur	<i>S₂</i>
	Cadmium	<i>Cd</i>		Tellurium	<i>Te₂</i>
	Silver	<i>Ag</i>	Tetra- Tri- atomic	Ozone	<i>O₃</i>
	Gold	<i>Au</i>		Phosphorus	<i>P₄</i>
	Lead	<i>Pb</i>		Arsenic	<i>As₄</i>
	Argon	<i>A</i>		Vanadium	<i>V₄</i>
	Helium	<i>He</i>			
	Krypton	<i>Kr</i>			
	Neon	<i>Ne</i>			
	Xenon	<i>Xe</i>			
	Nickel	<i>Ni</i>			
	Copper	<i>Cu</i>			
	Aluminum	<i>Al</i>			

and depend on the temperature at which the necessary data are obtained. Because of the variable atomicity of the sulfur molecule, its formula will be written simply as *S* in chemical equations representing reactions in which sulfur takes part. This phenomenon will be more fully discussed when the study of this element is taken up.

In writing chemical equations involving other elements, the usual atomicities of these elements, as given in Table I, will be employed throughout.

15. Standard for Determining Atomic and Molecular Weights.—Hydrogen was first used in determining atomic

and molecular weights, but, since oxygen combines with more substances than hydrogen, it has been thought best by chemists to use oxygen as the standard. In determining the atomic and molecular weights of the different elements, oxygen has been given a value of 16, twice its combining weight. By this standard the atomic weight of sulfur is 32.06. The atomic and molecular weights used by the chemist are relative, and not actual weights. For instance, the atomic weight of sulfur, which is 32.06, does not mean that sulfur has any particular weight; it simply indicates that the sulfur atom is approximately twice as heavy as an atom of oxygen. The weight of a molecule of oxygen is 32, since a molecule of oxygen contains 2 atoms (Table I), each with a weight of 16.

16. Equivalent and Atomic Weights.—The equivalent weight of an element when not identical with its atomic weight is a submultiple of the atomic weight. Thus, while the equivalent weight of oxygen referred to as a standard is 8, the atomic weight of this element referred to as a standard is 16. Therefore,

$$\text{Atomic weight} = n \times \text{equivalent weight}$$

Here n represents a small whole number, which indicates the valence of the element. For example, the atomic weight of oxygen is equal to its valence times its equivalent weight. The valence of oxygen is 2 and its equivalent weight is 8; therefore, its atomic weight is equal to 2×8 or 16.

In a similar manner the atomic weight of hydrogen is computed. Since the equivalent weight of hydrogen is 1.008 and the valence of hydrogen is 1, the atomic weight of hydrogen is 1×1.008 or 1.008.

The atomic weight of an element may be defined as the relative weight of an atom of an element as compared with the weight of one atom of oxygen taken as 16. The atomic weights of a number of different elements are given in Table II.

17. Determination of Molecular Weights.—The molecular weights of gases and of liquids and solids that can be vaporized without breaking the molecules can be determined by employing the principle that a liter of oxygen will have the same ratio to

TABLE II

INTERNATIONAL ATOMIC WEIGHTS, 1943

Reprinted from the Journal of the American Chemical Society

Element	Symbol	Atomic Number	Atomic Weight	Element	Symbol	Atomic Number	Atomic Weight
Actinium*	<i>Ac</i>	89	229.0	Mercury	<i>Hg</i>	80	200.61
Alabamine*	<i>Ab</i>	85	212.0	Molybdenum	<i>Mo</i>	42	95.95
Aluminum	<i>Al</i>	13	26.97	Neodymium	<i>Nd</i>	60	144.27
Antimony	<i>Sb</i>	51	121.76	Neon	<i>Ne</i>	10	20.183
Argon	<i>A</i>	18	39.944	Nickel	<i>Ni</i>	28	58.69
Arsenic	<i>As</i>	33	74.91	Nitrogen	<i>N</i>	7	14.008
Barium	<i>Ba</i>	56	137.36	Osmium	<i>Os</i>	76	190.2
Beryllium	<i>Be</i>	4	9.02	Oxygen	<i>O</i>	8	16.0
Bismuth	<i>Bi</i>	83	209.0	Palladium	<i>Pd</i>	46	106.7
Boron	<i>B</i>	5	10.82	Phosphorus	<i>P</i>	15	30.98
Bromine	<i>Br</i>	35	79.916	Platinum	<i>Pt</i>	78	195.23
Cadmium	<i>Cd</i>	48	112.41	Polonium*	<i>Po</i>	84	210.0
Calcium	<i>Ca</i>	20	40.08	Potassium	<i>K</i>	19	39.096
Carbon	<i>C</i>	6	12.01	Praseodymium	<i>Pr</i>	59	140.92
Cerium	<i>Ce</i>	58	140.13	Protactinium	<i>Pa</i>	91	231.0
Cesium	<i>Cs</i>	55	132.91	Radium	<i>Ra</i>	88	226.05
Chlorine	<i>Cl</i>	17	35.457	Radon	<i>Rn</i>	86	222.0
Chromium	<i>Cr</i>	24	52.01	Rhenium	<i>Re</i>	75	186.31
Cobalt	<i>Co</i>	27	58.94	Rhodium	<i>Rh</i>	45	102.91
Columbium	<i>Cb</i>	41	92.91	Rubidium	<i>Rb</i>	37	85.48
Copper	<i>Cu</i>	29	63.57	Ruthenium	<i>Ru</i>	44	101.7
Dysprosium	<i>Dy</i>	66	162.46	Samarium	<i>Sm</i>	62	150.43
Erbium	<i>Er</i>	68	167.2	Scandium	<i>Sc</i>	21	45.1
Europium	<i>Eu</i>	63	152.0	Selenium	<i>Se</i>	34	78.96
Fluorine	<i>F</i>	9	19.0	Silicon	<i>Si</i>	14	28.06
Gadolinium	<i>Gd</i>	64	156.9	Silver	<i>Ag</i>	47	107.88
Gallium	<i>Ga</i>	31	69.72	Sodium	<i>Na</i>	11	22.997
Germanium	<i>Ge</i>	32	72.6	Strontium	<i>Sr</i>	38	87.63
Gold	<i>Au</i>	79	197.2	Sulfur	<i>S</i>	16	32.06
Hafnium	<i>Hf</i>	72	178.6	Tantalum	<i>Ta</i>	73	180.88
Helium	<i>He</i>	2	4.003	Tellurium	<i>Te</i>	52	127.61
Holmium	<i>Ho</i>	67	164.94	Terbium	<i>Tb</i>	65	159.2
Hydrogen	<i>H</i>	1	1.008	Thallium	<i>Tl</i>	81	204.39
Illinium*	<i>Il</i>	61	146.0	Thorium	<i>Th</i>	90	232.12
Indium	<i>In</i>	49	114.76	Thulium	<i>Tm</i>	69	169.4
Iodine	<i>I</i>	53	126.92	Tin	<i>Sn</i>	50	118.7
Iridium	<i>Ir</i>	77	193.1	Titanium	<i>Ti</i>	22	47.9
Iron	<i>Fe</i>	26	55.85	Tungsten	<i>W</i>	74	183.92
Krypton	<i>Kr</i>	36	83.7	Uranium	<i>U</i>	92	238.07
Lanthanum	<i>La</i>	57	138.92	Vanadium	<i>V</i>	23	50.95
Lead	<i>Pb</i>	82	207.21	Virginium*	<i>Vi</i>	87	223.0
Lithium	<i>Li</i>	3	6.94	Xenon	<i>Xe</i>	54	131.3
Lutecium	<i>Lu</i>	71	174.99	Ytterbium	<i>Yb</i>	70	173.04
Magnesium	<i>Mg</i>	12	24.32	Yttrium	<i>Y</i>	39	88.92
Manganese	<i>Mn</i>	25	54.93	Zinc	<i>Zn</i>	30	65.38
Masurium*	<i>Ma</i>	43	97.8	Zirconium	<i>Zr</i>	40	91.22

* Atomic weight is not definitely established.

the weight of a liter of hydrogen or any gas as the weight of 1 molecule of oxygen will have to 1 molecule of hydrogen or any gas, under the same conditions of temperature and pressure. Knowing the weight of 1 liter of oxygen to be 1.43 grams, to calculate the molecular weight of hydrogen it is necessary to have only one additional fact, that is, the weight of 1 liter of hydrogen. The weight of 1 liter of hydrogen has been calculated by experiment to be 0.09009 gram. Having the weights of 1 liter of oxygen and 1 liter of hydrogen at standard conditions and knowing the weight of 1 molecule of oxygen to be 32, the molecular weight of hydrogen may be calculated by proportion, thus:

$$1.43 : 0.09009 :: 32 : x$$

$$x = 2.016 \text{ (molecular weight of hydrogen)}$$

It is found that sulfur dioxide gas weighs a trifle more than twice as much as an equal volume of oxygen. Therefore, since, according to Avogadro's hypothesis, equal volumes of gases contain the same number of molecules, the molecule of sulfur dioxide must weigh twice as much as the oxygen molecule. Since a molecule of oxygen has a weight of 32, the molecular weight of sulfur dioxide must be twice 32, or 64. The molecular weight of an element or a compound may be defined as the relative weight of a molecule of that element or compound as compared with the weight of a molecule of oxygen taken as 32.

18. For substances that are not gases nor liquids or solids that can be vaporized without breaking up the molecules, other methods must be used, which will not be described here. For substances that go into solution without decomposition, the molecular weight can be determined by its effect on the boiling and freezing points of the solution. Molecular weights may also be determined from the results obtained by accurate analysis of compounds made up in part of substances whose atomic and molecular weights are known.

If the number of atoms and the atomic weights of the elements in a molecule are known, the molecular weight of the compound can be easily calculated. In every case, the molecular weight of

a molecule is equal to the sum of its atomic weights. The molecular weight of any compound is therefore equal to the sum of the atomic weights as represented by the number of atoms in the molecule.

The method of calculating molecular weights from atomic weights should be known, since the molecular weights are used in every phase of chemistry and are indispensable to the industrial chemist. The application of Table II is shown by the following examples.

EXAMPLE 1.—Find the molecular weight of $NaCl$, sodium chloride.

SOLUTION.—The formula shows that 1 molecule of $NaCl$ contains 1 atom of each element. Therefore, the weights of these elements as given in Table II may be applied directly.

$$\begin{array}{rcl} \text{Atomic weight of } Na, \text{ Sodium,} & = & 22.997 \\ \text{Atomic weight of } Cl, \text{ Chlorine,} & = & 35.457 \\ \hline \text{Molecular weight of } NaCl & = & 58.454 \text{ Ans.} \end{array}$$

EXAMPLE 2.—Find the molecular weight of $CaCO_3$, calcium carbonate.

SOLUTION.—The formula shows that there are three atoms of oxygen, O . The weight given in Table II must therefore be multiplied by 3.

$$\begin{array}{rcl} \text{Atomic weight of } Ca, \text{ calcium,} & = & 40.08 \\ \text{Atomic weight of } C, \text{ carbon,} & = & 12.01 \\ 3 \times \text{atomic weight of } O, \text{ oxygen,} & = & 48.00 \\ \hline \text{Molecular weight of } CaCO_3 & = & 100.09 \text{ Ans.} \end{array}$$

EXAMPLE 3.—Find the molecular weight of $K_2Cr_2O_7$, potassium bichromate.

SOLUTION.—The formula shows that there are two atoms of K , two atoms of Cr , and 7 atoms of O ; therefore, the atomic weights of these elements must be multiplied by 2, 2, and 7, respectively.

$$\begin{array}{rcl} 2 \times \text{atomic weight of } K, \text{ potassium,} & = & 78.192 \\ 2 \times \text{atomic weight of } Cr, \text{ chromium,} & = & 104.020 \\ 7 \times \text{atomic weight of } O, \text{ oxygen,} & = & 112.000 \\ \hline \text{Molecular weight of } K_2Cr_2O_7 & = & 294.212 \text{ Ans.} \end{array}$$

EXAMPLES FOR PRACTICE

Find the molecular weight of:

- | | |
|---|--------------|
| 1. H_2SO_4 , sulfuric acid | Ans. 98.076 |
| 2. NH_4OH , ammonium hydroxide | Ans. 35.048 |
| 3. $K_3Fe(CN)_6$, potassium ferricyanide | Ans. 329.246 |
| 4. $Ba(OH)_2$, barium hydroxide | Ans. 171.376 |

19. Gram-Molecular Volume.—Frequently the chemist must use a weight, in grams, equal to the molecular weight of a substance. This weight is called the gram-molecular weight (G.M.W.) or molar weight. As the molecular weight of oxygen is 32, the gram-molecular weight, or molar weight, is taken as 32 grams.

The volume that the gram-molecular weight of a substance occupies is the gram-molecular volume (G.M.V.). By experiment it has been found that the gram-molecular weight of any gas under standard conditions will occupy approximately 22.4 liters. Since 1 liter of oxygen weighs 1.43 grams and the gram-molecular weight of oxygen is 32, it follows that 32 grams will occupy $\frac{32}{1.43}$ or approximately 22.4 liters.

If the gram-molecular weight of any gas is divided by its density, the result will be approximately 22.4 liters. It may, therefore, be concluded that the gram-molecular weight of a gas occupies 22.4 liters.

The gram-molecular volume is of great importance to the chemist, because it furnishes him with a very quick method for determining the molecular weight of any substance that can be weighed in the gaseous state. All the chemist need do is to compute the weight of 22.4 liters of the gas from the weight of a given volume of the gas.

EXAMPLE.—One liter of carbon dioxide weighs 1.977 grams. What is the molecular weight?

SOLUTION.— $22.4 \times 1.977 = 44$ grams. Ans.

METHODS OF DETERMINING ATOMIC WEIGHTS

20. Calculation of Atomic Weights from Molecular Weights.—The approximate atomic weights of elements that enter into gaseous compounds may be determined by the following steps:

1. Determine the molecular weight of a number of gaseous compounds that contain the element in question.
2. Determine the percentage, by weight, of the element in the compounds selected, by careful analysis.
3. Multiply the percentage by the molecular weight of the compound containing the element. The smallest product obtained is considered as the atomic weight of the element.

TABLE III
DETERMINATION OF ATOMIC WEIGHT OF NITROGEN

Name of Gaseous Compound	Formula	Molecular Weight	Percentage of Nitrogen by Analysis	Product of Percentage and Molecular Weight
Nitrogen	N_2	28.016	100.	28.016
Nitrous oxide	N_2O	44.016	63.65	28.016
Ammonia	NH_3	17.032	82.25	14.008
Nitrogen pentoxide..	N_2O_5	108.016	25.49	28.018
Nitrogen dioxide ...	NO_2	46.008	30.45	14.009

A concrete example will illustrate this method. Five gaseous compounds of nitrogen are listed in Table III, with their molecular weights; the percentage of nitrogen in them, determined by careful analysis; and the products obtained by multiplying the percentages of nitrogen in the compounds by the molecular weights of the compounds. The figures in the last column give the weight of nitrogen in the molecule. No compound of nitrogen has been found which has less than 14.008 grams of nitrogen in a gram-molecular weight. Since the atom is the smallest particle of an element that enters into a chemical reaction, it is therefore the smallest weight of an element that could be found

in a molecule. Hence, the atom of nitrogen must have an atomic weight of approximately 14.008.

21. Specific-Heat Method.—Dulong and Petit discovered by experiment that the product of the atomic weight of an element in the solid state and its specific heat equals a constant, 6.4.

$$\text{Atomic weight} \times \text{specific heat} = 6.4 \text{ (approx.)}$$

$$\text{Atomic weight} = \frac{6.4}{\text{Specific heat}} \text{ (approx.)}$$

The specific heat of a substance is the number of calories required to raise the temperature of 1 gram of the substance 1 degree centigrade. For water, the specific heat is 1; for most other substances it is less than 1. The elements listed in Table IV all have a specific heat of less than 1.

The specific-heat method is less exact than the previous method but it aids in deciding what multiple of the combining weight is the atomic weight.

The use of this method is illustrated as follows: The specific heat of sodium, Table IV, is .29, and this divided into 6.4 gives the approximate atomic weight of sodium as 22.07. This agrees very closely with the exact atomic weight of sodium, which is 22.997.

TABLE IV
SPECIFIC HEAT OF ELEMENTS

Element	Specific Heat	Element	Specific Heat
Lithium94	Iron112
Sodium29	Copper095
Magnesium245	Zinc093
Silicon16	Bromine (solid)084
Phosphorus (white)....	.20	Gold032

22. Combining Weights Method.—A more accurate method of determining atomic weights is by making use of the approximate atomic weights as determined by the specific-heat method or from the molecular weights and the combining weight, which

is determined with analytical precision. The atomic weight of sodium 22.997 (Table II) is equal to its combining weight, whereas, the atomic weight of oxygen is twice its combining weight. By these and similar observations, it may be concluded that the atomic weight of an element is equal to its combining weight or it is a multiple of its combining weight; that is,

Atomic weight = combining weight \times small whole number

To find the exact atomic weight, the following steps are followed:

1. Find the approximate atomic weight.
2. Determine the combining weight with analytical accuracy.
3. Multiply the combining weight by some small whole number that will yield a product nearest the approximate atomic weight. This product is the exact atomic weight.

From Table III it is impossible to say whether nitrogen has an atomic weight of 14.008 or 14.009. Since the combining weight of nitrogen has been determined as 14.008 by many accurate analyses, it can be definitely said that the product of this weight and some small whole number that equals the approximate atomic weight is the exact atomic weight. Hence, the exact atomic weight of nitrogen is $14.008 \times 1 = 14.008$.

SOLUTIONS

INTRODUCTION

23. Definitions.—A solution is a homogeneous mixture of two or more substances; it may be a solid, gaseous, or liquid mixture. Thus, a mixture of solids such as an alloy or glass is a solid solution; a mixture of gases such as we have in the atmosphere is a gaseous solution. Since the term solution is frequently restricted to liquid mixtures, we will, therefore, consider it as such here.

A solution may be defined as a homogeneous mixture of a liquid with a solid, a gas, or another liquid. The dissolving liquid is called the solvent, and the dissolved solid, gas, or liquid is called the solute. The mixture of solvent and solute is the solution.

In any true solution, intermingling of the particles of the separate components down to molecular magnitudes has actually been accomplished. In a solution of salt and water, the dissolved substance is completely and permanently dispersed throughout the solvent, water. No matter how long this solution is allowed to stand, the salt will never settle out. Complete separation can be effected only by evaporating all the water from the solution.

Practically speaking, there is no limit to the amount of dispersion which may be produced in a solution. A single crystal of potassium permanganate $KMnO_4$, which gives a deep purple-colored solution in water, may be dissolved in one liter or in 50 liters of water, and the purple color which it imparts to the solution will be as perfectly perceptible in every part of the solution. Here a very outstanding characteristic of a solution as opposed to a compound can be noted. The composition of a solution may be varied continuously within certain limits; whereas, compounds contain definite proportions by weight, and simple atomic ratios of their constituent elements.

24. Solvents.—The solvent may be, as stated before, a liquid other than water or it may be a solid or a gas. However, since water dissolves a greater number of substances than does any other known agent, aqueous solutions are the most common ones. Many organic substances, such as fats, rubber, tar, petroleum, paraffin, and shellac, do not dissolve to any measurable extent in water. Gasoline is a good solvent for fats, and for this reason a high-grade gasoline known as naphtha is used in dry-cleaning. Some organic substances, as sugar, dissolve easily in water, but hardly at all in other common solvents. Hence, candy or molasses can be taken out of clothing by water, but not by solvents of fats.

Fats dissolve readily in ether, $C_4H_{10}O$, in carbon bisulfide, CS_2 , in carbon tetrachloride, CCl_4 , and in chloroform, $CHCl_3$. This is the reason why these substances remove grease that accidentally gets into cloth. Tar and petroleum dissolve in gasoline and in benzene, C_6H_6 . Shellac is dissolved in alcohol to make varnish. Medicines that are dissolved in alcohol are called tinctures. Tincture of iodine is such a solution.

25. Suspensions, Colloidal Suspensions, and Emulsions.

If a very finely powdered substance such as flour is mixed with water, an apparently uniform, but cloudy, mixture results. The particles of flour are here simply suspended in the liquid, not dissolved, and will eventually settle out, the solid slowly settling to the bottom. Sand, shaken with water, settles at once. Such mixtures are called suspensions.

In some cases, while the subdivision of the suspended substance does not approach molecular magnitudes, it is so minute as to prohibit the substance from settling out in any reasonable time. Such suspensions are known as colloidal suspensions. Solutions of starch, glues, and gums are of this nature. The path of a beam of light, on passing through a suspension, appears milky, but on passing through a real solution it is not affected. The suspended particles are not in the form of single molecules, but are small aggregates or clumps of molecules.

If the substance is a liquid that is suspended throughout the solvent in little drops, the mixture is called an emulsion. Milk owes its cloudy, white appearance largely to droplets of oil matter (butter fat) suspended in water. They will easily pass through filter paper. When milk is allowed to stand, these droplets of oil matter slowly rise to the top, since they are lighter than the water in which they are suspended, but they are not dissolved.

The characteristics of a true solution, therefore, are homogeneity, absence of settling or separation, and extremely minute subdivision to molecular magnitudes of the dissolved substance.

VARIATIONS IN SOLUBILITY

26. Saturation.—The thought naturally arises as to whether there is any limit to the quantity of a substance that may be dissolved in a given quantity of solvent. As a rule, not more than a certain amount of solute is dissolved by a limited amount of solvent. If an excess amount of solute in a finely divided state is stirred and shaken with the solvent for a sufficient length of time, the maximum amount of solute will finally be dissolved. The solvent is then saturated with the solute. Thus, 100 ml. of water at 18° C. will dissolve as much as

35.86 g. of sodium chloride, NaCl , but no more. The same amount of water will dissolve 177.9 g. of sodium iodide, NaI , before the solvent becomes saturated. On the other hand, a saturated solution of calcium oxalate, CaC_2O_4 , in water will contain only 0.00055 g. in 100 ml. at 18°C . This particular case may be described by saying that sodium chloride is moderately soluble in water, sodium iodide very soluble, and calcium oxalate insoluble. No substance is absolutely insoluble, but substances like calcium oxalate are called insoluble because in most cases they may be so considered.

Many liquids will dissolve in one another in all proportions; in such cases there is no limit to the solubility, and therefore no possibility of the solution reaching saturation. Alcohol, glycerin, and water will dissolve in one another in any proportion. In such solutions the substances are said to be miscible in all proportions.

27. Solubility of Substances.—The number of grams of the solute required to saturate a fixed volume, say 100 ml., of the solvent is known as the solubility of the substance, at the existing temperature. Table V shows the solubilities of 142 substances in 100 ml. of water at 18°C . The table is divided into a number of horizontal and vertical rows. At the top of the vertical rows are placed a number of elements representing the positive radicals of the substances, and at the left of the horizontal rows are placed a number of elements or groups of elements representing the negative radicals. The upper number in the square at which two rows intersect shows the number of grams of the substance held in solution by 100 ml. of water at 18°C . The lower number is the molar solubility, that is, the number of mols contained in 1 liter of saturated solution. For instance, the square at which the rows headed K and Cl intersect contains the upper number 32.95, which represents the number of grams of potassium chloride, KCl , soluble in 100 ml. of water at 18°C .; and the lower number 3.9, which represents the number of mols (molecular-weights) of potassium chloride soluble in 1 liter of water at 18°C . The numbers for small solubilities have been abbreviated. Thus, $.064 = .0000004$.

TABLE V
SOLUBILITIES OF BASES AND SALTS IN WATER AT 18° C.

	K	Na	Li	Ag	Tl	Ba	Sr	Ca	Mg	Zn	Pb
Cl	32.95 3.9	35.86 5.42	77.79 13.3	.0813 .09	.3 .013	37.24 1.7	51.09 3.0	73.19 5.4	55.81 5.1	203.9 9.2	1.49 .05
Br	63.86 4.6	88.76 6.9	168.7 12.6	.01 .066	.042 .0015	103.6 2.9	96.52 3.4	143.3 5.2	103.1 4.6	478.2 9.8	.598 .02
I	137.5 6.0	177.9 8.1	161.5 8.6	.0835 .01	.006 .0817	201.4 3.8	169.2 3.9	200.0 4.8	148.2 4.1	419.0 6.9	.08 .002
F	92.56 12.4	4.44 1.06	.27 .11	195.4 13.5	72.05 3.0	.16 .0092	.012 .001	.0016 .0002	.0087 .0014	.005 .0005	.06 .002
NO ₃	30.34 2.6	83.97 7.4	71.43 7.3	213.4 8.4	8.91 .35	8.74 .33	66.27 2.7	121.8 5.2	74.31 4.0	117.8 4.7	51.66 1.4
ClO ₃	6.6 .52	97.16 6.4	313.4 15.3	12.25 .6	3.69 .13	35.42 1.1	174.9 4.6	179.3 5.3	126.4 4.7	183.9 5.3	150.6 3.16
BrO ₃	6.38	36.67	152.5	.59	.30	.8	30.0	85.17	42.86 *	58.43	1.3
IO ₃	7.62 .35	8.33 .4	8.2 3.84	.025 .0814	.009 .0016	.02 .001	.25 .0057	.23	6.87 .26	.83 .02	.002 .013
OH	142.9 18.0	116.4 21.0	12.04 5.0	.01 .001	40.04 1.76	3.7 .22	.77 .063	.17 .02	.001 .0002	.0005 .015	.01 .041
SO ₄	11.11 .62	16.83 1.15	35.64 2.8	.55 .02	4.74 .09	.0823 .01	.011 .0006	.20 .015	35.43 2.8	53.12 3.1	.0041 .0313
CrO ₄	63.1 2.7	61.21 3.30	111.6 6.5	.0025 .001	.006 .0001	.0835 .014	.12 .006	.4 .03	73.0 4.3065
C ₂ O ₄	30.27 1.6	3.34 .24	7.22 .69	.0034 .0317	1.48 .030	.0085 .0833	.0046 .0826	.0855 .043	.03 .0027	.0864 .04	.0316 .0854
CO ₃	108.0 5.9	19.39 1.8	1.3 .17	.003 .0001	4.95 .10	.0023 .0811	.0011 .07	.0013 .0313	.1 .01	.004 .083	.031 .033

28. Effect of Temperature on Solubility.—The quantity of a substance that will dissolve in a fixed amount of a given solvent depends largely on the temperature. The solvent action of liquids usually increases as the temperature rises, when the solute is a solid, and decreases as the temperature rises, when the solute is a gas. Thus, 100 g. of water will dissolve 13 g. of potassium nitrate, KNO_3 , at $0^\circ C.$ and 150 g. at $73^\circ C.$, an enormous increase in solubility. Cases in which the solubility of solids in liquids decreases with a rise in temperature are rare. The solubility of slaked lime, $Ca(OH)_2$, is .175 g. per 100 g. of water at $20^\circ C.$ and .078 g. per 100 g. of water at $100^\circ C.$

TABLE VI
SOLUBILITY OF GASES IN WATER

Name of Substance	By Volume			By Weight		
	$0^\circ C.$	$15^\circ C.$	$100^\circ C.$	$0^\circ C.$	$15^\circ C.$	$100^\circ C.$
Acetylene	1.73	1.15		.20	.13	
Air02881	.02045	.01105	.02936	.02090	
Ammonia	1298.9	802.4		98.7	60.6	
Carbon dioxide.	1.713	1.019		.3347	.1971	
Carbon monoxide03537	.02543	.0141	.0044	.0031	
Chlorine	4.61	2.635		1.46	.8493	
Hydrogen chloride	506.5	461.5		82.5	75.2	
Hydrogen0203	.0183	.1066	.00018	.00016	
Hydrogen sulfide	4.686	3.056		.710	.458	
Methane05563	.03690	.0170	.00396	.00260	
Nitrogen02388	.01786	.01	.00298	.00221	
Oxygen0489	.03415	.017	.00695	.0048	
Sulfur dioxide.	79.789	47.276		22.83	13.54	

Table VI shows the solubility of a number of gases in water at $0^\circ C.$, $15^\circ C.$, and $100^\circ C.$ The term "By Volume" in Table VI means the volume of the gases, at standard conditions, that may be dissolved by 1 volume of water at the different temperatures. The term "By Weight" means the number of

grams of gas, at standard conditions, dissolved by 100 grams of water at the different temperatures. For example, the table shows that .0489 volume of oxygen, at standard conditions, is dissolved in 1 volume of water at 0°C . to form a saturated solution, and that .00695 gram of oxygen, at standard conditions, is dissolved in 100 grams of water at 0°C . to form a saturated solution.

29. Supersaturated Solutions.—When a hot, saturated solution is cooled in the absence of any undissolved solute, there is some delay before the crystals of the excess solute begin to appear. This solution, until the appearance of the crystals, is said to be supersaturated. In most cases the crystals will appear in due course of time, especially if the supersaturated solution is stirred or shaken.

Some substances do not crystallize or separate from the solvent easily when a saturated solution is cooled. The compound sodium thiosulfate, commonly called "hypo," which photographers use in fixing negatives, is of such nature. If a hot, saturated solution of hypo is cooled, it may retain all of the hypo that dissolved in the water at the higher temperature. The cold solution contains more of the dissolved hypo than an equal volume of cold water would dissolve if placed in contact with hypo crystals. Such a solution is said to be supersaturated. If a small crystal of the hypo salt is placed in the supersaturated solution, the excess amount of solute will crystallize, leaving the exact amount of solute in solution that would actually dissolve in the water at that respective temperature.

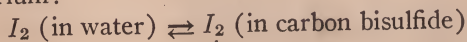
Crystallization may be started in a supersaturated solution by shaking or stirring the solution or by scratching the walls of the container with a sharp tool.

30. Henry's Law.—In 1803 Henry stated that at any definite temperature the solubility, by weight, of a gas in a liquid is proportional to the pressure of the gas. The volume of gas dissolved with an increase in pressure remains constant, even though the weight increases. At standard conditions (0°C . and 760 mm.) only 2 ml. of hydrogen will dissolve in 100 ml. of water, whereas at 2 atmospheres (1,520 mm.) 4 ml. of hydrogen

as measured at standard conditions will dissolve in 100 ml. of water. However, under the new pressure of 2 atmospheres, the 4 ml. of hydrogen will be compressed to 2 ml. You can readily see from the foregoing explanation that, although the weight of the hydrogen gas dissolved has increased, the actual volume of the gas dissolved is constant. Not all very soluble gases follow Henry's law, because of the fact that they react to some extent with water.

31. Law of Partition.—A very interesting application of the law of partition may be noted when two immiscible solvents are used. Iodine is very soluble in carbon bisulfide. A dilute solution of iodine in carbon bisulfide has a beautiful violet color. If a small particle of iodine is shaken with water, it will dissolve very slowly, eventually attaining a saturated but very dilute solution. If a sufficient quantity of carbon bisulfide is added and shaken with the aqueous solution of iodine, a violet color in the heavier carbon bisulfide layer will be observed. Most of the iodine has been removed from the aqueous solution by the carbon bisulfide, because it is more soluble in carbon bisulfide. The process of removing a substance from solution in one solvent and securing it in another solvent is called extraction.

In such cases, however, neither solvent can entirely deprive the other of all of the dissolved substance; a state of equilibrium is reached. In fact, the iodine is divided between the two solvents in the exact ratio of its solubility in the two solvents, if equal volumes of the solvents are used. For example, if the iodine is 100 times as soluble in carbon bisulfide as it is in water, only one part will remain in the water and 100 parts in the carbon bisulfide. The dissolved iodine is in constant motion, passing from one liquid to the other, as indicated by the following equilibrium:



The law of partition states that *the partition of the substance takes place in proportion to its solubility in each solvent.*

32. Concentrations of Solutions.—The concentrations of solutions, saturated or otherwise, are expressed in either physi-

cal or chemical units of weight. When physical units are used, the number of grams of the solute held in solution by a fixed quantity of solvent is given, usually 100 g. or 100 ml. of solvent.

When chemical units of weight are used, two methods are possible and both are used. Either the equivalent weights or molecular weights may be taken as a basis of measurement. The solutions based on equivalent weights are called normal solutions, and those based on molecular weights are called molar solutions.

33. A normal solution contains 1 gram-equivalent weight of the solute in 1 liter of solution. A gram-equivalent weight of a compound is that part of a compound which will interact with or replace 1 gram-equivalent weight of an element. In the discussion on valence it was stated that 1 atom of oxygen was equivalent to 2 atoms of hydrogen because it combined with them, and that 1 atom of chlorine was equivalent to 1 atom of hydrogen, since they combined atom for atom to form hydrogen chloride, HCl . From the foregoing statements it may be said that a gram-equivalent weight of a compound is that part of a compound which will combine with or displace 1 gram-equivalent weight of hydrogen (1.008 g.) or its equivalent. A molecular weight of hydrogen chloride, HCl (36.465 g.), is also a gram-equivalent weight, for it contains 1.008 g. of hydrogen, and this amount of hydrogen is replaceable by 1 gram-equivalent weight of a metal. The molecular weight of copper sulfate, $CuSO_4$, (159.63) contains 2 gram-equivalent weights, since the metal copper, Cu , has replaced 2 hydrogen atoms, H_2 , or 2 gram-equivalent weights of hydrogen (2×1.008 g.). Therefore, normal solutions of these two substances contain 36.465 g. of HCl and $\frac{159.63}{2}$ g. or 79.815 g. of $CuSO_4$ per liter of solution. The chief property of normal solutions is that equal volumes of different normal solutions contain the exact proportions of the solutes which are required for complete interaction.

A molar solution contains 1 gram-molecular weight (mol) of the solute in 1 liter of solution. In the cases previously cited, the molar solution contains 36.465 g. of HCl and 159.63 g.

of CuSO_4 in 1 liter of solution. It can be seen that the molar and normal solutions have the same concentrations when the radicals are univalent, as HCl .

34. Properties of Solutions.—It is a known fact that water vapor exerts a pressure which is directly proportional to the temperature. If, however, a solid substance is dissolved in water, the vapor pressure exerted by the solution is always less than that of the same amount of pure water at the same temperature. The lowering of the vapor pressure is directly proportional to the amount of solid substance that is dissolved. Because of this fact, the boiling point of such a solution must be higher than that of an equal amount of pure water under the same conditions.

When water boils at 100°C . it exerts a pressure of 760 mm. Since the vapor pressure of water increases with the rise in temperature, it can be readily seen that the temperature of a solution with a lower vapor pressure must be increased to a point where the solution will exert a vapor pressure of 760 mm. before it will boil. For example, at 100°C . a 7.5 per cent solution of potassium chloride, KCl , exerts a vapor pressure of 734.1 mm., while that of water at the same temperature is 760 mm. The difference is 25.9 mm. Therefore, the solution has to be raised to a higher temperature (100.96°C .) in order that a vapor pressure of 760 mm. is exerted before it boils. If the quantity of potassium chloride is doubled, the increase in the boiling point will also be doubled. In that case the solution will boil at 101.92°C .

35. Similar experience is encountered in connection with the freezing points of solutions. It is known that salt water does not freeze as easily as pure water. This is due to the fact that solutions freeze at a lower temperature than pure water. The amount of the lowering of the freezing point, just as in the case of the raising of the boiling point, has been found by experiment to be directly proportional to the concentration of the dissolved solid. When different solutes were investigated, the results obtained showed that the rise in the boiling point and the lowering of the freezing point depend on the molecular

concentration rather than on the weight of the dissolved substance.

For illustration, molecular weights of different substances were dissolved in equal volumes of water. Such solutions then contained the same number of molecules of the dissolved solute. The substances selected were cane sugar (mol. wt. = 342 g.), glucose (mol. wt. = 180 g.), phenol (mol. wt. = 94 g.), and urea (mol. wt. = 60 g.), all of which are very soluble in water. While the weight of solute in each case was different, the molecular concentration (molecules per unit volume) was the same. When the freezing points of the four different solutions were determined, they were all found to be very nearly the same, -1.86°C . In like manner the boiling point was found to be nearly the same for the four different solutions, $.52^{\circ}\text{C}$. above the temperature at which pure water boils. It is very evident, therefore, that the rise in the boiling point and the lowering of the freezing point depend on the number of molecules of dissolved solute rather than on its weight. It must be mentioned, however, that a relatively small group of compounds do not appear to obey this general rule.

IONIZATION

INTRODUCTION

36. Before ionization is discussed, it is very important to review the general nature of acids, bases, and salts. With a clear and better understanding of these three classes of substances, it will be easier to understand ionization. Although the properties of acids, bases, and salts differ materially from each other, the members of each class possess certain properties in common.

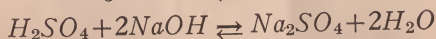
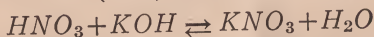
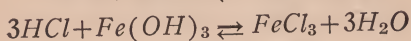
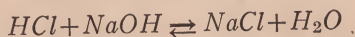
Acids are characterized by the ability to change the color of indicators, such as that of litmus from blue to red; they have a sour taste and react with most metals with the displacement of hydrogen.

Bases, on the other hand, are characterized by their ability to change the same indicators to a color different from that caused by an acid, such as that of litmus from red to blue; they possess a soapy feeling and an alkaline taste. Therefore, by means of

litmus it is possible to determine whether a solution is acidic or basic.

When solutions of acids, bases, and salts react with one another, they do so in a manner that seems to indicate that they are composed of two parts, called radicals. All bases contain a negative hydroxyl radical (OH^-) as one of these reacting parts. The positive radical is usually a single element, a metal. All acids contain a positive hydrogen radical (H^+) and a non-metal or group of non-metals acting as if it were a single element for a negative radical.

37. The following equations, representing reactions between acids and bases in solution, will make clear the nature of the radicals composing them:

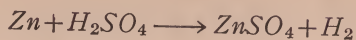


All these equations illustrate that an acid and a base will react with each other to form a salt and water. The hydrogen (H^+) of the acid and the hydroxyl (OH^-) of the base combine to form water, while the metal of the base and the non-metal radical of the acid combine to form a salt. When in solution, salts react with each other as if they were composed of the two radicals originally found in the acid and base. It must be noted that radicals have a definite valence or combining capacity, regardless of whether they exist as a part of an acid, a base, or a salt. The reaction that takes place between acids, bases, and salts is known as double decomposition. In double decomposition reactions, two compounds (acids, bases, or salts) react with each other to form two other compounds by exchange of radicals.

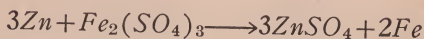
PROPERTIES COMMON TO ACIDS, BASES, AND SALTS IN SOLUTION

38. **Displacement.**—One of the radicals that belong to an acid, a base, or a salt in solution can be displaced by another element. The radical is obtained in the free state. For

example, hydrogen is liberated from acids by the addition of a metal :



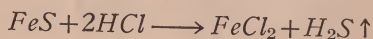
If a metal that is more active than the positive radical of the salt is added to a salt solution, it will displace the positive radical. For example, zinc will displace iron from a solution of ferric sulfate :



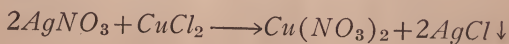
This principle is used in the purification of the more valuable metals. Thus, iron will displace gold from its various salt solutions.

39. Double Decomposition.—Examples of double decomposition reactions between acids, bases, and salts in solution have been discussed in Art. 37. Whenever two solutions of these compounds which do not contain a common radical are mixed, a double decomposition may result. Any base will react with an acid. Any acid or base may react with any salt of a different acid or base. Salts that do not contain a common radical will also react in pairs.

The important feature of a double decomposition is that, in all such reactions, each substance present behaves exactly as if it consists of two distinct radicals. Double decompositions may therefore be regarded as a result of the liberty of radicals to exchange partners. In some cases this exchange is incomplete, and the result obtained is an equilibrium mixture (reversible reaction), as illustrated in Art. 37. However, in many reactions, because one of the substances formed by the exchange of radicals may disappear from the solution either in the form of a gas or as a precipitate, the reaction practically goes to completion in one direction. For example :



The hydrogen sulfide, H_2S , leaves the reaction in the form of a gas. Also,



The silver chloride, $AgCl$, leaves the reaction to form a precipitate. The liberation of a gas produced in a chemical reaction is symbolized as \uparrow ; while the formation of a precipitate is symbolized as \downarrow .

Not only does every acid, base, and salt in solution behave in a double decomposition as if it consisted of two distinct radicals, but each radical possesses an independent set of properties. For example, the addition of the silver salt to the cupric chloride solution resulted in a precipitate of silver chloride. This is true of all soluble chlorides. If a soluble silver salt solution is added to a soluble chloride solution, a precipitate of silver chloride will form. A radical may therefore be defined as an atom, or group of atoms, which behaves as a distinct

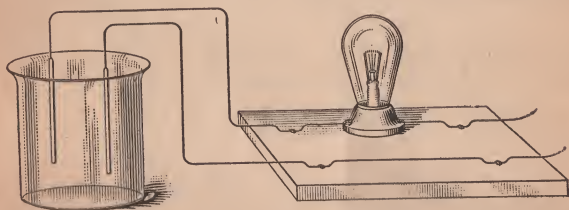


FIG. 2

unit in double decompositions, and which contributes a distinct and independent set of properties upon solutions of all acids, bases, and salts of which it forms one constituent.

40. Conductivity.—The water solutions of acids, bases, and salts are all conductors of electricity. For this reason they are classified as electrolytes. Solutions of substances which do not conduct electricity are classified as non-electrolytes.

When a current of electricity is passed through a solution of an acid, a base, or a salt, the solution will at all times decompose. A very careful study of this phenomenon has disclosed the fact that the radicals of which acids, bases, and salts are composed, always proceed toward the same pole under the influence of the electric current. For example, the hydrogen and metal radicals, which are positive radicals, always proceed toward the negative pole (cathode), and the hydroxyl and

other non-metal radicals, which are negative radicals, always proceed toward the positive pole (anode).

41. The conducting power of solutions may be examined by an apparatus, such as that in Fig. 2. The electrodes are connected to a direct-current circuit. The bulb, which is connected on one of the wires, shows by its glow when the solution is a conductor, and by varying brightness indicates the degree of conduction of the solution.

If the electrodes are placed in a solution of sugar in water, the bulb will not glow, because the solution will not conduct the current. However, if the electrodes are placed in solutions of acids, bases, and salts, the bulb will glow. By testing the conduction of different solutions of acids, bases, and salts, it is found that they do so in different degrees. Solutions of sodium and potassium hydroxides conduct well, and so do solutions of hydrochloric and nitric acids. Practically all salt solutions are good conductors. But there are many acids and bases, such as acetic acid and ammonium hydroxide, that are poor conductors. Solutions of acids, bases, and salts that are good conductors of electricity, are classified as strong electrolytes. Solutions that are relatively poor conductors are classified as weak electrolytes.

42. **Abnormal Boiling Points, Freezing Points, and Vapor Pressure.**—It has already been mentioned that a relatively small group of compounds did not act in accordance with the laws governing the raising of the boiling points, the lowering of the freezing points, and vapor pressure. Solutions of electrolytes give abnormal results. If an acid, a base, or a salt is dissolved in water and the solution is allowed to freeze, it will be found that the freezing point is lower than that of any other solution containing an equal molar concentration of solute. Likewise, the boiling point will be higher and the vapor pressure will be depressed to a greater degree.

The inconsistency of the abnormality with the theory that the lowering of the freezing points, the raising of the boiling points, and the depressing of the vapor pressure are directly proportional to the molar concentration of the solute, may be accounted

for by the theory of ionization. The theory of ionization assumes that the ions produce the same effect in lowering the freezing points, in raising the boiling points, and in depressing the vapor pressure as that produced by whole molecules.

43. Theory of Ionization.—The Swedish chemist Arrhenius proposed his theory of ionization in order to explain the phenomena of acids, bases, and salts discussed in the preceding articles, as well as many other of their characteristics. The theory may be summarized as follows:

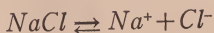
1. The molecules of many substances, when dissolved in water, dissociate into two or more parts called ions.

2. An ion differs from an atom or a molecule in that it carries an electric charge.

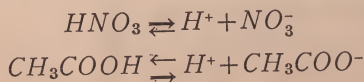
3. Two kinds of ions are formed: one charged with positive electricity and the other with negative electricity, the sum of the positive charges being sufficient to neutralize exactly all the negative charges.

4. Only electrolytes dissociate into ions.

The meaning of these laws may be more easily understood by considering the formation of a solution of common salt, sodium chloride, in water. According to the theory of ionization some of the molecules of sodium chloride, $NaCl$, dissociate when dissolved in water into ions of sodium, Na^+ , and chlorine, Cl^- . The dissociation of sodium chloride is shown by the equation



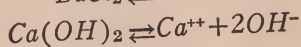
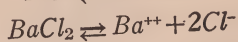
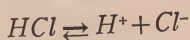
The foregoing equation also shows the equilibrium that exists between the undissociated and the dissociated parts of an electrolyte in solution. The degree of ionization of different electrolytes may be conveniently indicated by the thickness of the arrows. Thus, for nitric acid and acetic acid, we have the equilibria



In the case of the ionization of nitric acid the dark arrow indicates that there are more ionized (dissociated) parts in

solution than there are undissociated parts. The opposite is true of the solution of acetic acid.

44. Since the solution of hydrochloric acid or of any electrolyte, has no charge, there must be equal quantities of positive and negative electricity present. Therefore, it follows that univalent radicals become ions carrying a single charge, bivalent radicals become ions carrying double charges, trivalent radicals become ions carrying triple charges, etc. For example:



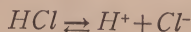
In the foregoing equations the charges of the Cl^- and OH^- ions are multiplied by the coefficients before them. A study of these reactions will show that there are equal numbers of positive and negative charges present. In the case of $BaCl_2$ there is one Ba^{++} ion with two positive charges present and two Cl^- ions, each with one negative charge present, and in the case of $FeCl_3$ there is one Fe^{+++} ion with three positive charges present and three Cl^- ions, each with one negative charge present. Therefore, all ions bear quantities of electricity in proportion to their valence. An ion may therefore be defined as an atom or group of atoms bearing an electric charge, or a number of charges equal to its valence. A positive ion is a free atom or group of atoms that has lost an electron or a number of electrons; while a negative ion is a free atom or group of atoms that has gained an electron or number of electrons. The positive ions are called cations because of the fact that they are attracted to the negative electrode, called cathode, while the negative ions are called anions because they are attracted to the positive electrode, called anode.

45. **Nomenclature.**—Because of the fact that the ionized elements or radicals are entirely different in both physical and chemical properties from their corresponding free elements or radicals, they must be given separate names. The representative names in Table VII are used for the ionic substances.

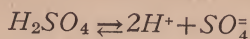
TABLE VII
IONIC NAMES OF RADICALS

Symbol	Name of Substance	Cation of Salt	Symbol	Name of Substance	Anion of
Na^+	Sodium-ion	Sodium	OH^-	Hydroxyl-ion	Bases
K^+	Potassium-ion	Potassium	$SO_4^{=}$	Sulfate-ion	Sulfates
Cu^{+++}	Cupric-ion	Cupric Copper	HSO_4^-	Hydrosulfate-ion	Bisulfates
Cu^{++}	Cuprous-ion	Cuprous Copper	Br^-	Bromide-ion	Bromides
NH_4^+	Ammonium-ion	Ammonium	Cl^-	Chloride-ion	Chlorides
H^+	Hydrogen-ion	Acids	$CO_3^{=}$	Carbonate-ion	Carbonates
Fe^{++}	Ferrous-ion	Ferrous iron	NO_3^-	Nitrate-ion	Nitrates
Fe^{+++}	Ferric-ion	Ferric iron	ClO_4^-	Chlorate-ion	Chlorates
Ag^+	Silver-ion	Silver	CH_3COO^-	Acetate-ion	Acetates

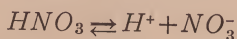
46. Ionization of Acids.—When an acid is dissolved in water, it dissociates into hydrogen-ions, which always serve as cations, and into ions of the remaining radical, which always serve as anions. Hydrochloric acid dissociates upon solution into hydrogen-ions and chloride ions. Thus:



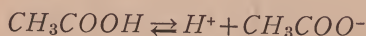
Sulfuric acid dissociates upon solution into hydrogen-ions and sulfate-ions:



Nitric acid dissociates upon solution into hydrogen-ions and nitrate ions:



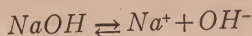
Acetic acid dissociates upon solution into hydrogen-ions and acetate ions:



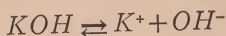
In like manner, other acids form cations of hydrogen. Therefore, an acid may also be defined as a substance that produces hydrogen ions, when dissolved in water.

47. Ionization of Bases.—When a base is dissolved in water it dissociates into a positive metal ion and a negative hydroxyl-ion.

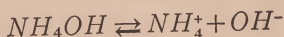
Sodium hydroxide, when dissolved in water, dissociates into sodium-ions and hydroxyl-ions:



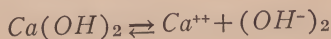
Potassium hydroxide, when dissolved in water, dissociates into potassium-ions and hydroxyl-ions:



Ammonium hydroxide, when dissolved in water, dissociates into ammonium-ions and hydroxyl-ions:



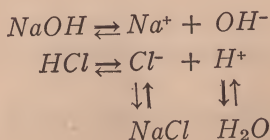
Calcium hydroxide, when dissolved in water, dissociates into calcium-ions and hydroxyl-ions:



In like manner, other bases produce hydroxyl-ions when dissolved in water. Therefore, a base may also be defined as a substance that produces hydroxyl-ions, when dissolved in water.

48. Ionization and Salts.—When an acid and a base are mixed together in the proper proportions the properties of each substance disappear. The resulting solution tastes neither sour nor bitter but is, on the contrary, salty. It has no effect on litmus nor on other indicators.

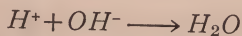
A study of the change that takes place when a solution of hydrochloric acid is combined with the proper proportions of a solution of sodium hydroxide, shows that water is always one of the products formed. The other product is a salt. Written in ionic form the change that takes place is shown by the following equation:



The heavy arrows indicate, in each case, the direction in which the equilibria tend to become established. The HCl , $NaOH$, and $NaCl$ are all very highly ionized, while the ioniza-

tion of the water is very slight. When the H^+ ion and OH^- ion meet, they combine to form a molecule of water, which does not re-ionize to any appreciable extent. This reaction removes the H^+ ion and OH^- ion from the field of action, thus causing the equilibria to shift to the right with a final result that nearly all of the H^+ ions and OH^- ions are used up.

49. The reaction shows that water is formed, and, unlike sodium and chlorine, does not remain dissociated as hydrogen and hydroxyl ions while in solution. The Na^+ and Cl^- , however, remain as free ions in the solution, and, strictly speaking, no sodium chloride is formed until the water is evaporated out of the solution. A study of reactions of this kind has shown that hydrogen and hydroxyl ions cannot exist together in solution to any extent, but combine to form molecules of water:



The change that forms a neutral substance when an acid reacts with a base in the proper proportions is called neutralization, and, in the light of the theory of ionization, may be defined as the union of the hydrogen-ion of an acid with the hydroxyl-ion of a base to form water. In like manner, a salt is formed by the union of the cation of a base with the anion of an acid. In the foregoing example, the cation (sodium) of the base sodium hydroxide combines with the anion (chlorine) of the hydrochloric acid to form the salt sodium chloride, $NaCl$.

50. **Degree of Ionization.**—If the assumption, that the ions produce the same effect on the freezing point, boiling point, and vapor pressure as that produced by whole molecules is correct, the degree of ionization of an acid, a base, or a salt may be determined by determining the freezing or boiling points of their solutions.

For example, it is known that the freezing points of solutions of sugar, glucose, phenol, urea or any other non-electrolyte, containing 1 molecular weight of solute in 1 liter of solution, is $-1.86^\circ C$. Therefore, if a solution of common salt, $NaCl$, containing 1 molecular weight of solute per 1 liter of solution is

prepared, this solution should freeze at the same temperature, -1.86°C ., provided the molecules of NaCl do not ionize. On the other hand, if all the molecules of NaCl ionize, the freezing point should be -3.72°C ., since each molecule would yield two ions, (Na^+ , Cl^-). A solution of NaCl , however, freezes at -3.50°C . The difference in freezing point, $3.50-1.86$, or 1.64° , is due to the large number of free ions in solution. The degree of ionization is therefore $\frac{1.64^{\circ}}{1.86^{\circ}}=88$ per cent.

51. Degree of ionization is also calculated by electrical conduction. Since an electrical current passed through a solution is conducted only by the ions present, it follows that the amount of current carried is directly proportional to the number of ions present in the solution.

Table VIII lists a number of the more common acids, bases, and salts, showing the approximate degree of ionization, as calculated by freezing determinations and conductivity determinations, in tenth-normal water solutions at 18°C . In the case of acids or bases containing more than one displaceable hydrogen or hydroxyl ion, the kind of ionization on which the per cent of ionization is based should be noted.

A study of Table VIII will show that the per cent of ionization of water is extremely low, that salts are relatively very highly ionized, and that acids and bases ionize with a considerable variation. A study of the table will also show that the more active acids and bases are highly ionized, which indicates that the ions are the active agents in chemical changes involving these substances. The standard by which the chemist determines the strength of acids and bases is by the degree of ionization. The strong acids and bases show a high degree of ionization when in solution, while the weaker acids and bases show a low degree of ionization when in solution.

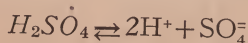
52. **Ions and Electrolysis.**—A solution of sulfuric acid in water is assumed to contain three kinds of solute: undissociated H_2SO_4 molecules, positively charged H^+ ions, and negatively charged SO_4^- ions. These are scattered indiscriminately throughout the solution. Equilibrium is maintained by a con-

TABLE VIII

DEGREE OF IONIZATION IN TENTH-NORMAL SOLUTION AT 18° C.

Acids	Per Cent	Acids	Per Cent
Hydrochloric acid (H^+ , Cl^-)	92	Acetic acid (H^+ , CH_3COO^-)	1.3
Nitric acid (H^+ , NO_3^-)	92	Carbonic acid	
Sulfuric acid ($2H^+$, SO_4^{2-})	61	(H^+ , HCO_3^-)	0.17
Oxalic acid (H^+ , $HC_2O_4^-$)	50	Hydrosulfuric acid	
Phosphoric acid		(H^+ , HS^-)	0.07
(H^+ , $H_2PO_4^-$)	27	Boric acid (H^+ , $H_2BO_3^-$)	0.01
Hydrofluoric acid (H^+ , F^-)	15	Water (H^+ , OH^-)	0.0001
Bases	Per Cent	Bases	Per Cent
Sodium hydroxide		Barium hydroxide	
(Na^+ , OH^-)	91	(Ba^{++} , $2OH^-$)	77.
Potassium hydroxide		Ammonium hydroxide	
(K^+ , OH^-)	91	(NH_4^+ , OH^-)	1.3
Calcium hydroxide		Water (H^+ , OH^-)	0.0001
(Ca^{++} , $2OH^-$)	90		
Salts	Per Cent	Salts	Per Cent
Potassium chloride		Barium chloride	
(K^+ , Cl^-)	86	(Ba^{++} , $2Cl^-$)	76
Sodium chloride (Na^+ , Cl^-)	86	Potassium sulfate	
Ammonium chloride		($2K^+$, SO_4^{2-})	72
(NH_4^+ , Cl^-)	85	Silver sulfate ($2Ag^+$, SO_4^{2-})	70
Potassium nitrate (K^+ , NO_3^-)	83	Zinc sulfate (Zn^{++} , SO_4^{2-})	40
Sodium acetate		Copper sulfate	
(Na^+ , CH_3COO^-)	79	(Cu^{++} , SO_4^{2-})	40
Sodium bicarbonate		Mercuric chloride	
(Na^+ , HCO_3^-)	78	(Hg^{++} , $2Cl^-$)	1

tinual dissociation of the H_2SO_4 and recombination of H^+ ions and SO_4^{2-} ions, as shown by the following equation:



The general method used in the electrolysis, or electrical decomposition, of a solution is illustrated in Fig. 3. Two plates a and

b, called electrodes, made of suitable materials (carbon or metals) are connected to the source of an electric current, such as an electric cell *c*, and dipped into the electrolyte *d*. The electrode *b* connected to the negative pole of the battery is the negative electrode, or cathode, while the electrode *a* connected to the positive pole is the positive electrode, or anode.

As soon as the circuit is completed and the current enters the solution, all of the ions in solution begin to migrate toward their proper electrodes. The positive and negative ions are attracted to the opposite electrodes. The positively charged H^+ ions, called cations, are attracted toward the negative electrode, or the cathode; while the negatively charged SO_4^- ions,

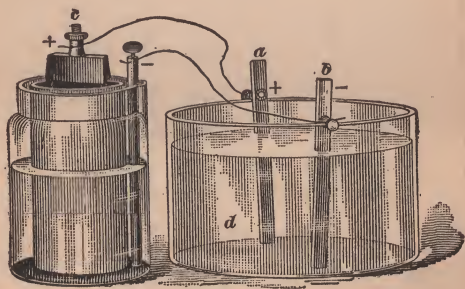
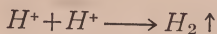


FIG. 3

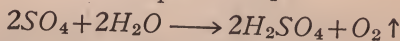
called anions, are attracted toward the positive electrode, or the anode. In all cases the solution is decomposed by electrolysis. Unless special conditions prohibit, the radicals or ions are liberated at the electrodes.

53. When sulfuric acid is electrolyzed, its hydrogen is liberated at the negative electrode, while oxygen is liberated at the positive electrode. In order to understand the action that takes place, it must be remembered, as previously mentioned, that the ions of the elements differ altogether from the elements themselves, inasmuch as they possess electrical charges which profoundly alter their characters. In order that an ion may appear as an element, its electrical charge must be removed, and this is exactly what happens during electrolysis. In the case of sulfuric acid, the positively charged H^+ ions are attracted

to the negatively charged cathode, and, when they come into contact with the cathode, they give up the charge and become electrically neutral atoms. Two such neutral atoms unite to form a molecule of free hydrogen, which is liberated at the cathode, thus:

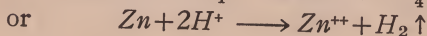
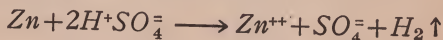


On the other hand, the negatively charged SO_4^- ions are attracted to the positively charged anode. When they come in contact with the anode, they give up their charge to become electrically neutral sulfate ions. These ions in turn react with water to form sulfuric acid and oxygen, which is liberated at the anode. The reactions that take place are:

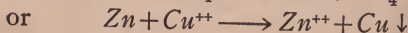


The sulfuric acid of the foregoing equation, a secondary product of electrolysis, is itself electrolyzed, and so on indefinitely.

54. Ionic Displacement.—When a metal acts upon an acid solution and hydrogen is liberated, it is the ions which are directly responsible for the action. For example the action of zinc on sulfuric acid is written:



Zinc has a similar action on solutions of salts, the metal of which is below zinc in the electromotive series. Zinc displaces copper from a solution of copper sulfate:



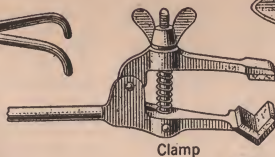
All metals can be arranged in such an order that each metal displaces the metals that follow it and is in turn displaced by the metals that precede it. This arrangement of metals is known as the electromotive series of metals, because, in the electrolysis

ELECTROMOTIVE
SERIES OF
METALS

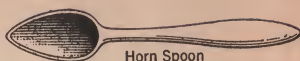
Potassium
Sodium
Barium
Calcium
Magnesium
Aluminum
Manganese
Zinc
Chromium
Iron
Cadmium
Cobalt
Nickel
Tin
Hydrogen
Antimony
Bismuth
Arsenic
Copper
Mercury
Silver
Platinum
Gold



Tongs



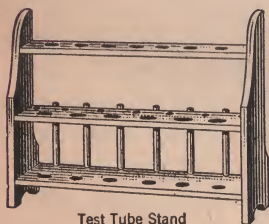
Clamp



Horn Spoon



Water Bath



Test Tube Stand



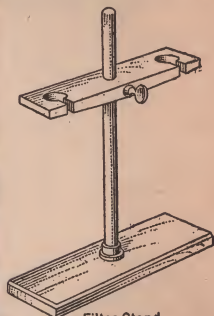
Separatory
Funnel



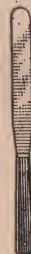
Pipette



Hydrometer



Filter Stand



Spatula



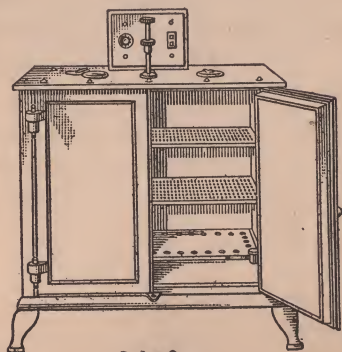
Forceps



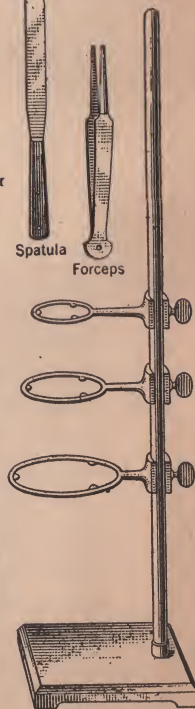
Tripod



Fish-Tail Burner



Drying Oven



Ring Stand and Rings

of normal solutions of the metallic ions, the electromotive force of the current required to deposit each metal is less than that for the preceding metal. The list represents the order in which the free elements tend to lose electrons. The higher the metal is in the list, the more readily will its atoms enter the ionic state, giving up or transferring electrons to positive ions of metals below it in the series.

The electromotive series furnishes a key to all actions involving solutions in which a free metal is used or produced.

PHYSICAL PROCESSES

APPARATUS AND METHODS

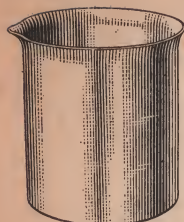
PRINCIPAL TYPES OF APPARATUS

55. Good work can be done by a chemist only when he is thoroughly familiar with the names and uses of the various kinds of apparatus needed to perform experiments and to make tests. The different types of apparatus shown on Plates I, II, and III will be described in detail in connection with the many operations dealt with in different parts of the text. It is advisable, however, to learn to associate its name with the appearance of each piece of apparatus illustrated, for it is knowledge that will be helpful in practical laboratory work. For this purpose each illustration is provided with the name of the instrument, apparatus, tool, etc., that it is intended to represent.

It should be understood, however, that the apparatus here shown represents but a small portion of that required by a commercial laboratory engaged in making all kinds of tests. In most lines of commercial laboratory work, apparatus designed especially for certain kinds of work is often used.

METHODS OF SEPARATION

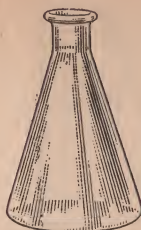
56. **Classification of Methods.**—There are some operations in laboratory work that are constantly applied, particularly in the separation of substances from one another. These methods are of such general use that it is well for the student of chemistry to know what details must be observed and what points must be



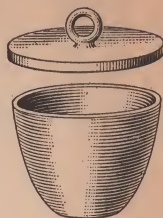
Beaker



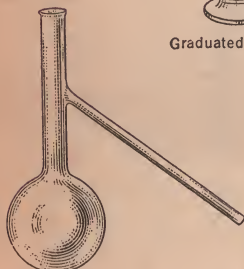
Graduated Cylinder



Erlenmeyer
Flask



Crucible
and Cover



Distilling Flask



Woulff's Bottle



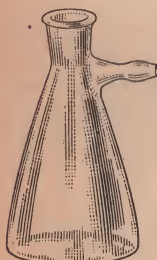
Wash Bottle



Round Bottom
Flask



Watch Glass



Filter Flask



Weighing Bottle



Wash Bottle



Florence Flask



Sand Bath



Evaporating Dish



PLATE III

noted in them. In each case, advantage is taken of the fact that a constituent to be separated from a mixture has some property not possessed by the rest of the constituents.

The principal methods of separation are: Solution, decantation, filtration, crystallization, evaporation, distillation, sublimation, and ignition. The details to be observed in each of these processes should be studied carefully, for they find constant application in chemical work and are indispensable in laboratory operations. In this part of the text, only the details of operation are discussed. The principles upon which they are based are explained in other lessons.

57. Solution, Decantation, and Evaporation.—If common table salt is placed in water it soon disappears and leaves a clear liquid mixture. The salt is said to have dissolved in the solvent, water, to form a solution of salt in water. Increasing the surface area of a substance by powdering it and heating the solvent after adding the substance to it, hastens the solution. A mortar and pestle are generally used to powder substances.

The fact that a substance is soluble enables it, under certain circumstances, to be separated from other substances. Suppose, for example, that a mixture of sand and salt is to be separated into its constituents. It has been determined that sand is insoluble (cannot be dissolved) in water and that salt is soluble (can be dissolved): facts that enable one to separate the substances completely.

The mixture is shaken up with water, a process that dissolves the salt and leaves the sand unaffected. The sand is allowed to settle to the bottom, and the clear solution on top, called the supernatant liquid, is carefully poured off into a basin without disturbing the settled sand. The operation of pouring off a supernatant liquid without disturbing the sediment is called decantation. Washing by decantation is accomplished by pouring off the supernatant liquid, adding fresh water, mixing thoroughly, allowing the insoluble substance to settle, and again pouring off the supernatant liquid.

If heat is applied to the vessel containing the salt solution, water passes off as steam and crystals of salt form in the basin.

The conversion of a liquid to a vapor is called evaporation, a process that is used extensively in chemical operations.

Either a test tube or a beaker can be used as the vessel in which salt, water, and sand are mixed; an evaporating dish serves as the vessel in which to collect the salt solution and from which to evaporate the water. Heat can be applied by means of either a Bunsen burner or an alcohol lamp. A piece of wire gauze is generally placed on a tripod, and the basin containing the salt solution is placed on the gauze.

58. Filtration.—Unlike the sand mentioned in the foregoing article, there are some solids so finely divided that they float in a liquid and settle only with difficulty, in which case the process of decantation cannot be used. Filtration, the opera-

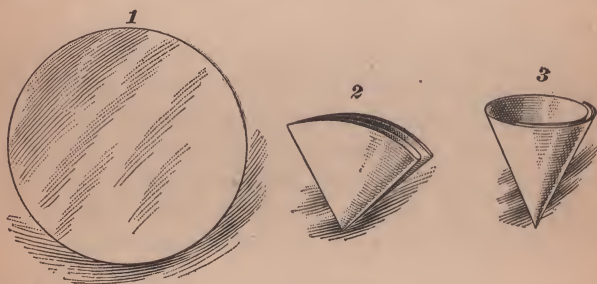


FIG. 4

tion that must be used in this instance, is the process of separating a liquid from an insoluble substance by pouring the mixture on a filter paper, a porous material, in which case the liquid passes through the pores and the insoluble substance remains on the filter paper.

For illustration, suppose that a mixture consisting of powdered chalk and salt is to be separated into its constituents. The mixture is thoroughly shaken with water to dissolve the salt and is then poured on a folded filter paper, arranged in a glass funnel. The solution of salt will pass through the filter paper and the chalk will remain on the filter. The liquid that passes through the porous material during filtration is called the filtrate. Asbestos, talc, and cotton are porous materials often used for filtering purposes. In every case, one should wet a

filter paper after it has been placed in a funnel and before any of the material to be filtered is introduced.

59. In Fig. 4 is shown how a filter paper is folded for use in a funnel. A piece of soft porous paper, as shown at 1 (blotting paper 3 inches in diameter will do for ordinary practice), is doubled twice and then opened into a cone, taking three folds of the paper on the one side and one on the other, as shown at 2 and 3 respectively. The cone of paper is placed in a glass funnel and moistened with water. Care should be taken that the point of the folded filter is not broken.

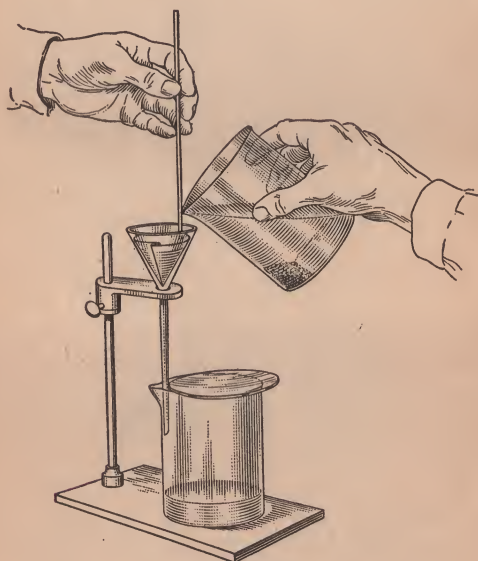


FIG. 5

As filtration is one of the most important of all processes used in the laboratory, one should observe carefully every detail upon which the operation depends. As an illustration, let it be supposed that a solution of sodium carbonate is added to a solution of barium chloride. A white insoluble substance is formed, as a result of the chemical reaction which takes place:



This insoluble substance, barium carbonate, $BaCO_3$, is called a precipitate, which is defined as an insoluble substance formed from a solution. Precipitation, or the process of forming a precipitate, is brought about, as a general rule, by the formation of an insoluble substance, as a result of a chemical action between two soluble substances in solution.

Let it be supposed that the precipitate formed when the solution of sodium carbonate is added to the solution of barium chloride, is to be separated from the rest of the mixture. The

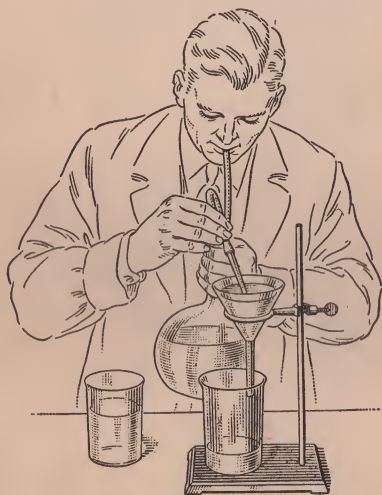


FIG. 6

solution and precipitate are poured on the filter by holding a glass rod or thin piece of glass tubing against the lip of the beaker or test tube, and by tipping the vessel containing the mixture, as shown in Fig. 5. The clear liquid, or filtrate, passes through the pores of the filter paper and is collected in a suitable vessel such as a beaker, the precipitate remaining on the filter. Some clean water is poured on the filter paper and the precipitate in order to wash them free from any soluble substances that have not yet passed through the paper.

Care should be taken, when pouring the mixture on the filter, to hold the glass rod or thin glass tubing close to the filter paper

in order to guard against spattering and to prevent the liquid from running down on the outside of the vessel. A wash bottle is generally used, as shown in Fig. 6, to supply the water needed to wash the soluble substances through the filter paper. The wash bottle is also used, as shown in Fig. 7, to remove the last portion of the mixture from the vessel containing it.

60. Crystallization.—A mixture may contain two soluble substances that differ from each other in their degree of solubility. In cases such as this, the mixture is completely dissolved

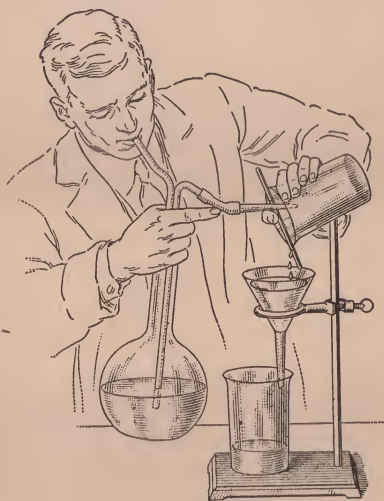


FIG. 7

by boiling. As the solution cools, the less soluble substances will separate out in a definite shape and the more soluble substances will remain in solution. The process in which a soluble substance separates from a solution, as a definitely shaped body, is called crystallization, and the definitely shaped body is called a crystal.

For example, a mixture of equal parts of potassium chlorate and potassium chloride is placed in a test tube, and just enough water is added to dissolve the substances upon boiling. On allowing the solution to cool, crystals of the less soluble potassium chlorate will separate from the solution. These crystals

may be removed from the solution by filtration, then purified by washing with a small amount of cold water, and dried at a very low temperature.

61. Fractional Crystallization.—If equal parts of sodium and potassium nitrates are dissolved in hot water, and the solution is then slowly cooled, potassium nitrate will crystallize first and then the sodium nitrate. Each lot of crystals may be redissolved and the process of cooling and crystallization repeated, each repetition of the process resulting in the further purification of the two salts. This is called fractional crystallization.

62. Distillation.—In the case of a mixture of a liquid and a solid or a mixture of two or more liquids, it is sometimes necessary to separate the component parts of the mixture into the original constituents; in this case the process of distillation is used. Distillation is a process, in which a substance is first liquefied, if it is a solid, by means of heat, or if already a liquid, it is evaporated, or volatilized, by the application of heat to the containing vessel, the volatile matter, or vapor, then being passed through a condenser and again liquefied or condensed.

For example, if it is desired to separate a mixture containing colored matter and water (a water-ink mixture will answer), the work may be done as follows: The mixture is placed in the round-bottomed distilling flask *a*, Fig. 8, which is connected to the condenser *b* and contains the thermometer *d*. Heat is applied to the flask by means of the modified Bunsen burner *h*, the flask being supported by means of an iron ring covered with a piece of wire gauze and clamped to the ring stand *i*. The flask is held in its vertical position by means of the universal clamp *j* and the clamp holder *k*. When the liquid in the flask begins to boil or distil, the vapors pass through the side tube of the flask into the condenser, which is supported by the ring stand *e*, the universal clamp *l*, and the clamp holder *m*. This condenser is kept cold by water that enters the water-jacket, not the tube itself, at *f* and leaves at *g*. When the vapors strike the cold inner tube of the condenser, they liquefy or condense and are collected in the flat-bottomed Florence flask *c*, which in this

case acts as the receiver. The condensed liquid, termed the condensate or distillate, is the water of the original mixture. It will be found to be perfectly clear, the coloring matter of the ink being left behind as a residue in the distilling flask.

63. Fractional Distillation.—Fractional distillation may be illustrated as follows: Suppose there is a solution of benzene

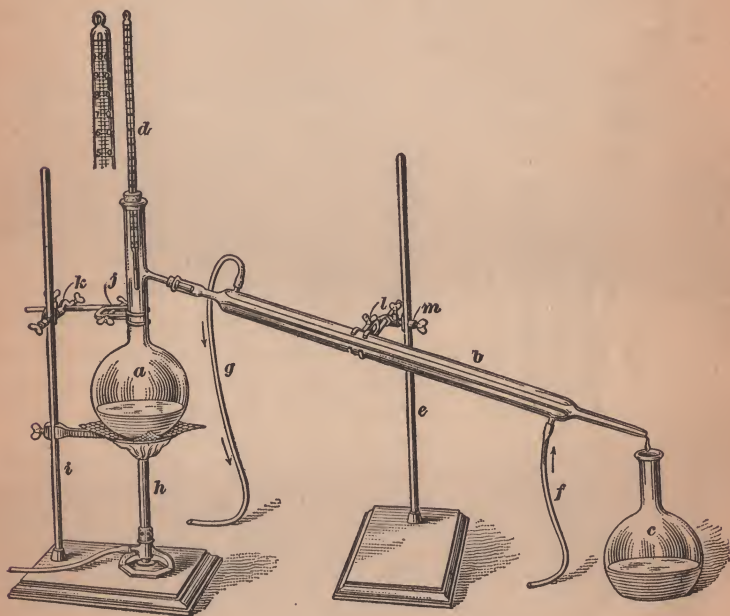


FIG. 8

and xylene, two organic coal-tar products, and that it is required to separate them as completely as possible by means of distillation. The boiling point of benzene is 80°C ., while that of xylene is about 142°C . Taking 100 milliliters of a mixture containing 50 milliliters of each, and distilling it, the greater part of the benzene will come over first, while, as the distillation proceeds, the distillate will become richer and richer in xylene. Then, if in the first distillation two portions, or fractions, of 50 milliliters each had been collected, the first would have contained most of the benzene and the second most of the xylene. If these

fractions are again separately distilled and other fractions collected at 10°C . below and above the boiling points of benzene in the one case and of xylene in the other, it will be seen that, if this procedure is continued for a sufficient number of distillations, two fractions will ultimately be obtained, which will consist of nearly pure benzene and xylene. Of course, results as good as these can hardly be obtained with the ordinary laboratory apparatus, but they may be obtained by employing specially designed distilling apparatus.

64. Distillation Under Diminished Pressure.—Throughout the subject of chemistry expressions such as the following are often found: *distillation under diminished pressure*, *distillation in vacuo*, and as it is important that the meaning of these expressions should be understood, they are here explained.

In Fig. 9 is shown a manometer of the form commonly used in laboratories when making distillations at pressures less than that of the atmosphere. At sea level the average pressure exerted by the atmosphere is equal to 14.7 pounds per square inch, and this pressure will support a column of mercury 760 millimeters in height, in a vertical, exhausted tube.

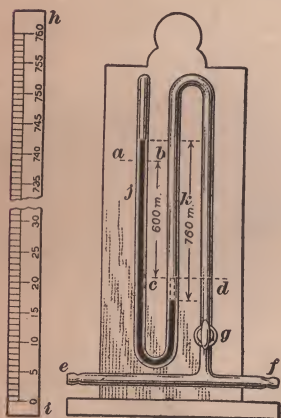


FIG. 9

65. The apparatus illustrated in Fig. 9 consists of a bent glass tube, sealed at the upper end and fitted with a stop-cock *g* at the other end; this end is joined to a short tube *e f*. If the ends *e, f* and the stop-cock *g* are left open, the pressure exerted upon the lower end of the mercury column is that of the atmosphere. When a vacuum pump is attached to the end *f* after the end *e* has been closed, the distance between the two levels of the mercury column, normally equal to 760 millimeters, will gradually decrease as the pump is operated. The levels may shortly occupy the positions indicated by the lines *a b* and *c d*, which correspond to a pressure of 600 millimeters.

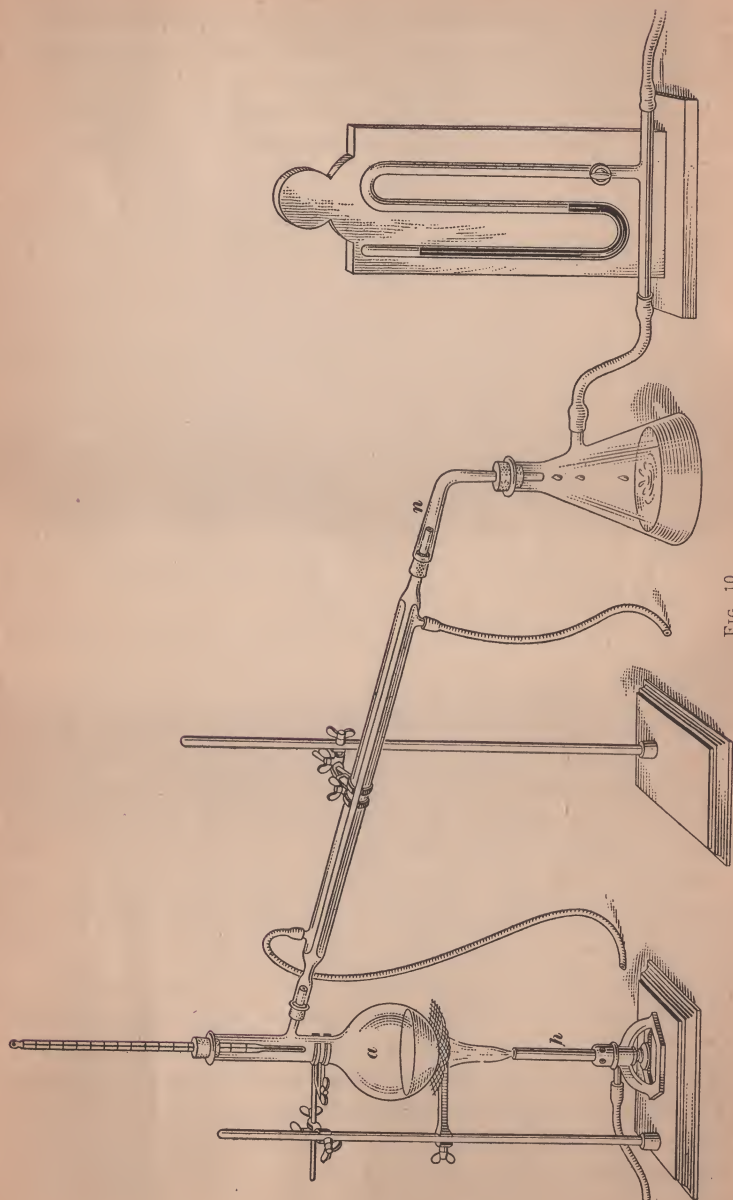


Fig. 10

Continuing the operation of the vacuum pump, the levels of the mercury columns in the two limbs *j* and *k* tend to approach each other as zero pressure is reached.

In commercial instruments, a scale graduated in millimeters, similar to that shown at *h i*, is employed. This scale may be adjusted so that its zero point is on the same level as that of the lower end of the mercury column, thus allowing the distance between the levels of the mercury in each limb of the manometer to be read off directly on the scale.

66. When a distillation under diminished pressure is to be made, the apparatus may be set up as shown in Fig. 10. One of the open ends of the manometer is connected to some type of vacuum pump, while the other is connected to the side-tube of an ordinary filter flask, which acts as a receiver. These connections are made by means of heavy, thick-walled rubber tubing, commonly called pressure tubing. For the sake of convenience, an adapter *n* is used between the lower end of the condenser and the top of the filter flask.

In low-pressure work, it is very necessary that all joints be tight or it will be impossible, in most cases, to obtain the decrease in pressure desired and, for this reason, all of the stoppers used as shown in Fig. 10 are made of rubber. However, if rubber is attacked by the substance being distilled, cork stoppers which have been coated with collodion may be used. Collodion is a solution of nitrocellulose, the solvent usually being a mixture of ether and alcohol.

Having set up the apparatus as shown and with the substance to be distilled in the flask *a*, to begin the distillation it is merely necessary to light the burner *h*, start the water going through the condenser, and then start the vacuum pump.

Distillations under diminished pressure are usually resorted to when organic compounds are to be distilled which decompose when heated to their boiling points at atmospheric pressure. Glycerine, for example, boils at 290° C., but it rapidly decomposes at this temperature; therefore, in order to obtain the practically pure colorless product, it must be distilled under diminished pressure, whereby the boiling point is reduced.

67. Sublimation.—When heated, most solids, as, for example, ice, will first form a liquid and then a vapor. There are a few solids, however, such as iodine, ammonium chloride, camphor, etc., which pass directly from a solid form to that of a vapor when heated, without first forming liquids. The distillation of substances that condense directly into solids without forming liquids, is called sublimation, and the condensed substances are called sublimates.

The process of sublimation is sometimes used to separate substances from one another. For example, if a mixture of sand and ammonium chloride is placed in a test tube and then heated, dense white fumes of ammonium chloride are given off which, upon cooling, condense in the upper part of the tube as a white sublimate. The sand in the mixture is unaffected.

Substances that can be distilled, evaporated, or sublimed are called volatile, and those that cannot be are called non-volatile bodies.

68. Ignition.—Some substances burn and form gases and others do not. This fact affords another means of separating substances that form a mixture. The substance that burns and generally forms a gas, called a combustible substance, passes from the mixture and leaves the substance that does not burn, called an incombustible substance, behind. The process of heating a substance intensely is called ignition.

As an example, suppose that charcoal and sand are to be separated from each other. If this mixture is heated to redness, the pure charcoal (carbon), a combustible substance, burns and forms a gas, carbon dioxide, which passes from the mixture; while the sand, an incombustible substance, remains behind.

69. Catalysis.—Catalysis is the acceleration or retardation of the speed of a chemical reaction by the presence of a comparatively small amount of foreign substance which does not alter in quantity or chemical composition during the reaction. Thus, hydrogen peroxide may be decomposed into water and oxygen by means of gold or silver, neither of which is altered in the least in the process of decomposition, but each exerts a catalytic action.

70. Effervescence, Deliquescence, Efflorescence.—Effervescence is the escape of a gas from a liquid in which it is either generated or has been held by pressure, as carbonic acid obtained from dissolving marble or from soda water. Deliquescence is the dissolving of a substance in water which it absorbs from the moisture present in the air. Examples are sodium and potassium hydroxides. Such substances are said to be hygroscopic, although not all hygroscopic substances dissolve in the absorbed water. Efflorescence is the gradual crumbling of a crystal to powder on exposure to the air, due to the escape of the water necessary for the formation of the crystal. Examples are: copper sulfate, $CuSO_4 \cdot 5H_2O$, and ferrous sulfate, $FeSO_4 \cdot 7H_2O$.

METHOD OF WEIGHING

71. Analytical Balance.—As a knowledge of the correct manipulation of an analytical balance is of importance to the chemist, a general description of the balance and its operation is given at this point.

The balance illustrated in Fig. 11 (*a*) has a capacity of 200 grams and is sensitive to a weight of one-twentieth of a milligram. The balance is enclosed in a glass case fitted with sliding doors in front and in back. The drawer in the base contains the weights, brushes, spatula, etc. At the top of the hollow pillar *a* is embedded a rectangular-shaped piece *b* of agate, on which rest the agate knife edges *c* of the beam *d*. The stirrups *e*, from which the scale pans *f* are suspended, also rest on agate knife edges. At *g* is shown the leveling screw by which the case may be adjusted so as to bring the air bubble in the spirit level *h* into a central position.

The knob *i* operates a rod, placed in the pillar *a*, by means of which the beam may be lifted slightly off its bearings, or knife edges. This is the position to be occupied by the beam when the balance is not in use. By turning the knob *i* in the required direction, in this case to the left, the beam is lowered upon its knife edges and the balance is ready for weighing operations.

72. As the materials to be weighed have a more or less corrosive action on metals, the pans *f*, Fig. 11, carry balanced

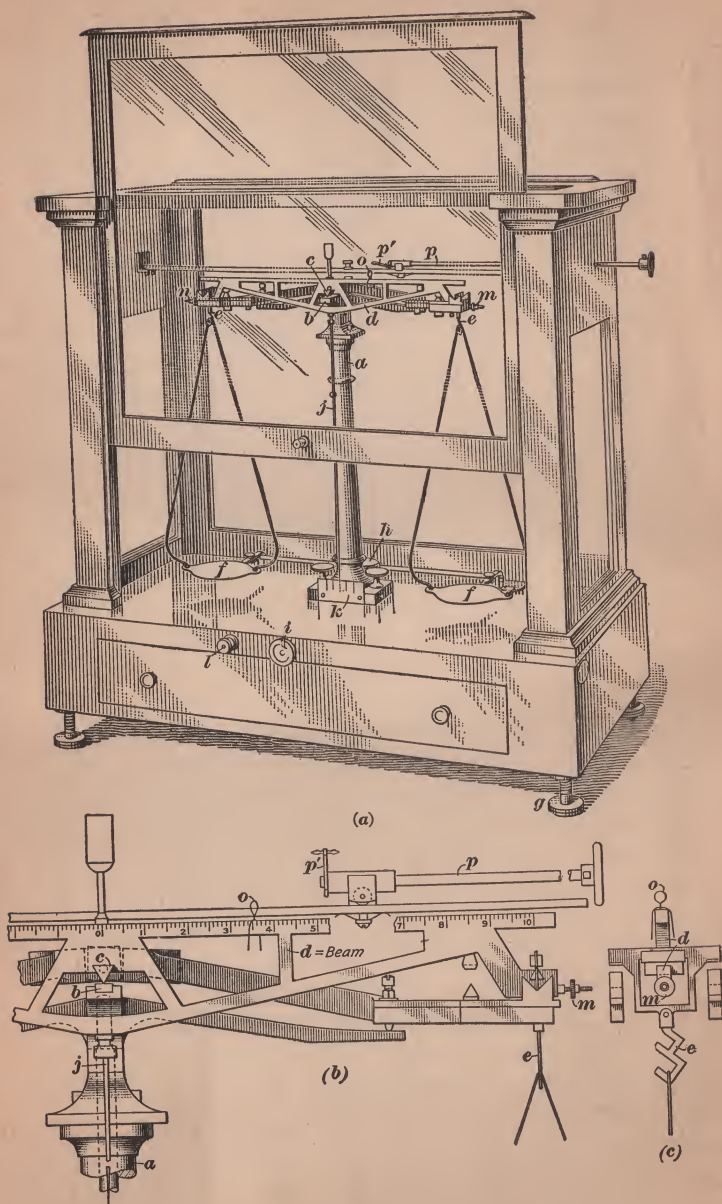


FIG. 11

watch glasses on which the chemicals are placed. However, if noncorrosive solids are to be weighed, the pan at the left is provided with an aluminum scoop-shaped pan the weight of which is exactly balanced by a special weight placed on the other pan. It is important to ascertain if the two pans are in perfect balance, in which case the pointer, or index, j should come to rest at the zero mark of the scale k , when the beam is released.

How far this is the case may be tested in the following manner: The balance is released by turning the knob i to the left; also, the scale pans are released by pushing in the knob l and turning it to the right. The beam and the pans are now free to swing and if the pointer moves over the same number of scale divisions on either side of the zero mark it indicates that the pans are in balance.

Should the pointer swing farther to the left of the zero mark, it proves that the right-hand pan is too heavy and that balance must be established by means of the screws m and n . In this case the screw m would be turned so as to bring it nearer the pillar a ; or, the screw n would be turned in the same direction, thus bringing it farther away from a . Either of the screws should be given only a few turns at a time, testing the results and repeating the operations until the pointer swings over the same number of divisions on either side of the zero mark, coming finally to rest at this point.

Fig. 11 (b) is an enlarged view of the right-hand end of the beam partly broken away. In this view the rider o is clearly shown; also, the arm p' on the rod p by which the rider is moved into position. At the end of the beam is seen the adjusting screw m . At (c) is an end view of the beam, corresponding parts in all the views being indicated by similar letters.

73. The operation of weighing consists in placing the substance to be weighed on the left-hand watch glass and the required number of weights on the other glass to produce a balance between the two pans, thus compelling the pointer to come to rest at the zero mark. The smallest separate weight that may be used for the right-hand pan weighs 10 milligrams. In order to weigh closer than this, it is necessary to employ

the platinum rider *o*, shown resting on the beam *d*, Fig. 11. This rider weighs 10 milligrams, and is known as a 10-milligram rider. When placed on the beam at the division marked 10, it produces the same effect as a 10-milligram weight placed on the right-hand pan. If this rider is moved to the division 1 of the beam or to any other division up to 10, the effect produced corresponds with that obtained by placing a correspond-

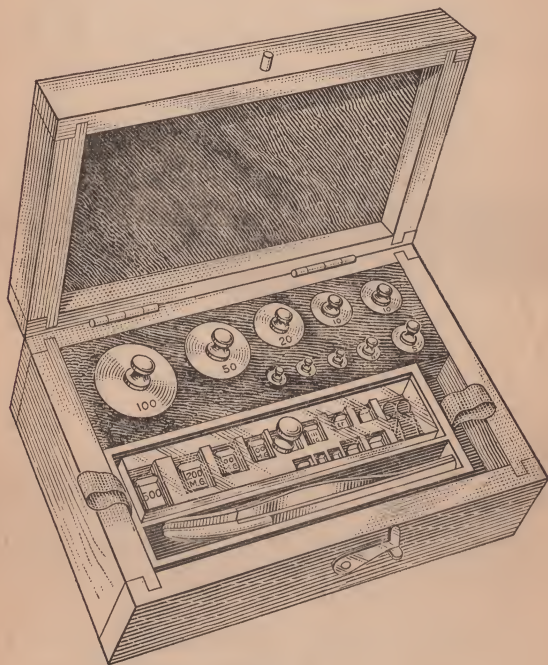


FIG. 12

ing number of milligrams on the right-hand pan. In some balances there is a rider for each end of the beam and either one may be placed into position by means of the rods *p*.

74. The weights are usually kept in a box, as shown in Fig. 12. The larger weights are generally made of brass and sometimes gold-plated; the smaller ones are made of aluminum or platinum. Each small weight rests in a separate compartment under a glass lid.

Accurate weights must under no consideration be touched with the fingers, but always lifted with the forceps.

The arrangements of the weights varies with different makes. Fig. 12 shows the usual and most convenient arrangement for a set of weights of from 100 grams to .001 gram.

Before attempting to weigh, the student must learn not only the denomination of each weight, but also its place in the box. He must be quite as well able to read the weights he has placed in the balance pan from the empty space in the box as from the weights themselves.

PRACTICAL HINTS

PHYSICAL OPERATIONS

75. Heating Glass.—Glass is one of the most useful of all materials to the chemist in his work, for it is transparent, can be put into different shapes when hot and plastic, and is unaffected by most chemicals. It is, however, brittle, a property that makes it liable to break, unless handled with care.

Heat should never be applied to glass unevenly lest the glass expand unevenly and crack. A wire gauze is used when heating liquids in glass vessels to keep the flame from coming into direct contact with the glass, and to secure an even distribution of heat. In this way, equal expansion of all parts of the glass is obtained and there is little danger that the glass vessels will crack.

76. Bending and Drawing Glass Tubing.—Heat will make glass so soft that it can be bent into various shapes. Hold in each hand one end of the glass tube to be bent, and place it lengthwise in a flat flame (generally produced by a fishtail burner), as shown in Fig. 13; do not use the direct flame of the Bunsen burner for this work, for it will cause the glass to crease at the bend. Keep turning the glass slowly in the flame at the point where you wish to bend it until it softens and begins to sag. Then raise the ends, as shown in Fig. 14, until the desired bend is obtained. Care should be taken to soften the glass thoroughly before bending it; precautions should also

be taken to keep turning the tube in the flame lest heat be applied to only one side, with the result that the bend will be flat instead of even and round.

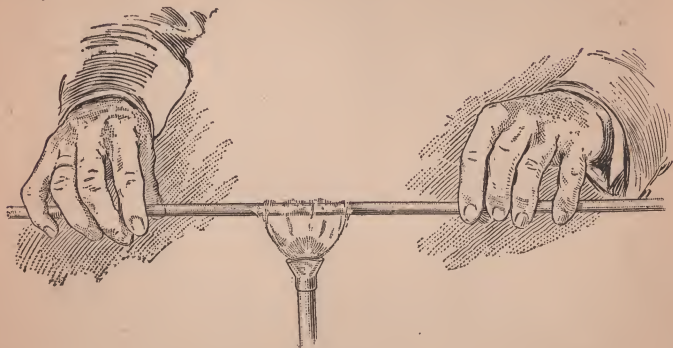


FIG. 13

77. In order to make a capillary tube, as at *BA*, Fig. 15, a piece of glass tubing is heated at a middle point as in the foregoing operation. When it has softened, remove the glass

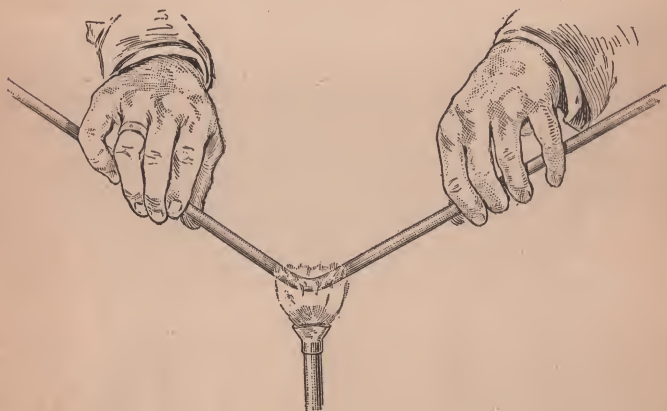


FIG. 14

from the flame and draw the ends quickly in opposite directions. The tube will now appear as shown in Fig. 15, and the part of the tube between *A* and *B*, which has very thin walls and a small bore, may easily be broken out. This part of the tube

is called a capillary tube and is used mostly for making melting-point determinations, as will be shown later.

78. Cutting Glass Tubing.—A chemist often needs to cut a piece of glass tubing to a desired length, a task that is easily accomplished. Scratch the glass tube with a sharp, triangular file and then hold the tube in both hands; the two forefingers should be together immediately under the scratch and the two



FIG. 15

thumbs above. Turn the wrists outward and downward, as though trying to widen the nick in the tubing, until the tube breaks. Fig. 16 illustrates this operation.

79. Rounding Ends of Glass Tubing.—It is often necessary to round the sharp end of a piece of glass tubing so that it may be passed through rubber or cork stoppers without difficulty and inserted into rubber tubing without cutting the inside walls of the latter. The rounding is best accomplished by holding the sharp end in a Bunsen-burner flame until it begins to soften, at which time it becomes smooth and round. A

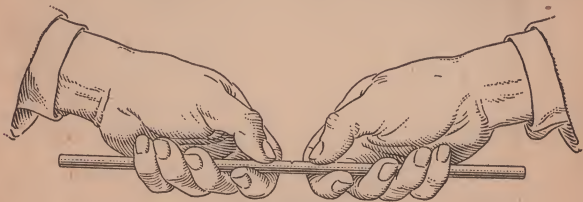


FIG. 16

moistened end will pass into a stopper or rubber tube more readily than a dry one. For this reason, a little vaseline is sometimes rubbed around a tube end or it is dipped into water before it is inserted.

80. Boring Corks and Rubber Stoppers.—When apparatus is set up in the laboratory, it is generally necessary to use corks and rubber stoppers, and since these are sold in the solid form, unless otherwise specified, they must be bored in order

to be used as shown in Figs. 8 and 10. For this purpose an instrument called a corkborer, Fig. 17, is used; it consists of a set of several thin metal (usually brass) tubes, the lower ends of which are sharpened. A rod is also contained in the set, in order to thrust out of the borer the piece of cork cut from the holes. The rod can also be pushed through two holes in the top of the borer, forming a handle to grasp when boring. Before commencing to bore, see that the borer is clear; then, beginning at the small end of the cork, bore carefully by twisting the borer and pushing it through. Take

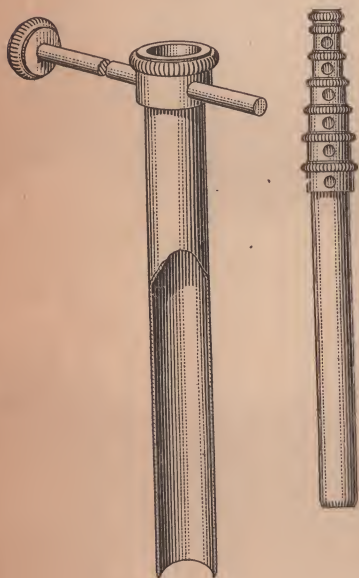


FIG. 17

care, while this is going on, that the hole is perfectly straight. In order to insure a clean-cut hole at the point where the borer emerges from the cork, the latter may be placed over a block of soft wood so that, when the borer comes through, the cork will be cut cleanly, while, because of the soft wood, the instrument will not be dulled. The borer selected for the work should be just a trifle less in diameter than the glass tubing to be used. This will tend to produce an air-tight fitting. It will be found of advantage to wet the glass tube with water,

in order to facilitate its passage through the hole in the cork. While one is attempting to force the glass tube into this hole, the hand should be covered with a cloth as a protection in case the tube should break. Remember that the success of the experiment depends, to a large degree, on the fact that the stopper, as well as the glass tube, fits air-tight. Much time and trouble is saved by setting up every piece of apparatus as nearly perfect as possible.



INORGANIC CHEMISTRY

(PART 2)

Serial 5560B

Edition 1

EXAMINATION QUESTIONS

Notice to Students.—*Study the Instruction Paper thoroughly before you attempt to answer these questions. Read each question carefully and be sure you understand it; then write the best answer you can. When your answers are completed, examine them closely, correct all the errors you can find, and see that every question is answered; then mail your work to us.*

Note.—Show how you obtain answers to questions that involve the solution of problems.

(1) (a) State the law of conservation of mass. (b) Show by a chemical reaction the application of the law.

(2) Using Table II, calculate the molecular weights of the following: (a) Potassium dichromate, $K_2Cr_2O_7$; (b) Nickelous ammonium sulfate, $NiSO_4 \cdot (NH_4)_2SO_4$; (c) Perchloric acid, $HClO_4$; (d) Mercuric arsenate, $Hg_3(AsO_4)_2$.

(3) Summarize the theory of ionization.

(4) Define the following and give an example of each; (a) solution; (b) emulsion; (c) solute; (d) solvent; (e) suspension.

(5) A certain gaseous substance is found upon analysis to be composed of the element oxygen and no other element; 500 ml. at standard temperature and pressure are found to weigh 0.7150 gram. What is the molecular weight of the gas?

(6) Define: (a) catalysis; (b) effervescence; (c) deliquescence; (d) efflorescence; (e) sublimation.

(7) Calculate the degree of ionization of hydrochloric acid in a solution containing one gram-molecular weight of HCl in 1 liter of solution, whose freezing point is $-3.60^\circ C$.

(8) (a) State the law of definite proportions. Illustrate by example. (b) State the law of multiple proportions. Illustrate by example.

(9) Using the table of Specific Heats, determine the approximate atomic weights of the following elements: (a) lithium; (b) zinc; (c) gold; (d) iron; (e) magnesium.

(10) (a) Define thermochemistry. (b) What is a thermochemical equation? Give an example of a thermochemical equation.

(11) (a) What weight of each of the following substances is contained in 1 liter of normal solution: (1) NaCl ; (2) H_2SO_4 ; (3) FeCl_3 . (b) What is the difference between a normal and a molar solution?

(12) Give two conditions that will cause a double decomposition reaction to go to completion. Illustrate by example.

(13) (a) Give six methods by which substances are separated from one another in a chemical laboratory. (b) Define filtration. Give an example of filtration.

(14) Calculate the weight of 1 liter of each of the following gases at standard temperature and pressure.

GAS	MOLECULAR WEIGHT
NH_3	17.032
H_2S	34.076
NO_2	46.008

(15) Write an ionic equation representing the ionization of each of the following: (a) acid; (b) base; (c) salt.

Mail your work on this lesson as soon as you have finished it and looked it over carefully. DO NOT HOLD IT until another lesson is ready.

